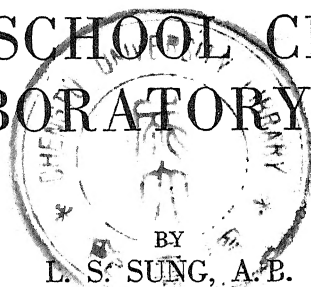


# MIDDLE SCHOOL CHEMISTRY AND LABORATORY MANUAL



ST. JOHN'S UNIVERSITY, SHANGHAI

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1930



## FOREWORD

It is difficult to write for the beginner a suitable laboratory manual because chemistry is an unfamiliar subject. If the directions of the laboratory manual are too brief, leaving the student to guess what is to be done, frequent failure will result, thus destroying the student's confidence in himself. On the other hand, if the directions are so very complete that the results to be obtained are too evident, requiring no observation, no conclusion, no serious thinking, and no further effort from outside reading by the student, the consequence will be a training worse than superficial.

This manual is the outgrowth of a number of years of practical teaching by the writer. The experiments, requiring only simple apparatus and common chemicals, have been tried and modified to suit the mental attitude of the beginner. At the beginning of the experiments there is an admirable summary of facts and theories of the experiments following, giving the student a proper amount of introductory information, and immediately after the summary is a large number of suggestive and interesting questions, the solution of which calls for a liberal amount of careful outside reading. I am sure that teachers using this book will find it a well-organized and well-balanced manual.

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*Head of Chemistry Department,  
Nankai University,  
Tientsin.*

AUG 10, 1928.





## PREFACE

At the present time there is a great need for a fundamental knowledge of the science of chemistry by our young people. It will open their eyes not only to the elements of a wonderful science, but will give them also an idea of the abundance of natural resources at the disposal of the people of our country—a country which abounds in mineral wealth and raw products which, if developed in the right way, will make this country one of the richest in the world.

A proper study of chemistry, which is one of the most useful sciences, will help the young student to learn system, acquire accuracy, stimulate interest in natural phenomena, encourage more careful observation of his surroundings, and stimulate the desire to be more independent in thought and action.

The following piece of work is a selection of facts and theories from many different sources and from the experiences and observations made on different classes in middle school chemistry. The material has been selected carefully, and an attempt has been made to make the subject matter as simple and concise as possible and readily understood by students of Chinese senior middle schools. Special emphasis has been laid on the presentation of fundamental facts and theories—the groundwork on which the science is built. Everyday facts related to chemistry have been included in the hope that they will serve to arouse the desire to investigate further our daily contacts. Thought-provoking questions are presented at the end of each chapter to lead the student to do his own thinking. It is the sincere hope that more of our young people will learn to think and work independently of their teachers—therefore the inclusion of these questions. One of the greatest values of the study of a scientific subject is the awakening of the desire to investigate, and to seek truth, and to do that the student must learn to apply himself to his task.

Enough material has been presented to allow for variation and selection from year to year.

Special acknowledgment is made to the authors of "Elementary Principles of Chemistry," and "Laboratory Exercises to Accompany Elementary Principles of Chemistry," Raymond B. Brownlee, Robert W. Fuller, William J. Hancock, Michael D. Sohon, and Jesse E. Whitsit, for the inspiration of this volume, for the headings of the chapters, and for most of the exercises of the "Laboratory Manual"—which have been selected from their manual and simplified for the use of Chinese middle school students. Acknowledgment is also made to the authors of the texts listed in the Bibliography for material used; also to the *National Geographic Magazine* for some of the illustrations, to the *China Journal of Arts and Science* for illustrations 31, 32, and 33, and to the *Saturday Evening Post* for illustration 43.

L. S. SUNG.

ST. JOHN'S UNIVERSITY,  
SHANGHAI, CHINA,  
JUNE, 1928.

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# MIDDLE SCHOOL CHEMISTRY AND LABORATORY MANUAL

## CHAPTER I INTRODUCTION

Chemistry is one of the most fascinating of all subjects to study. Many natural phenomena which formerly were regarded with superstition and uncertainty have been explained by chemical laws. Man is surrounded by the forces of nature and must deal with them constantly, and these same chemical laws have helped him to overcome many obstacles and difficulties. A knowledge of the elementary laws of nature will not only aid in exploring the world of reality widely and deeply but will also aid us in understanding those citizenship problems which arise in connection with such subjects as the use of waste products, pure foods, elimination of smoke and dust, purification of water, etc.

There are elements of real value for social and moral training found in studying the lives of the men who have built up the science of chemistry. From their biographies we can find inspiration and guidance. At the age of nineteen years, Perkin started a factory for making dyes, showing at that early age courage, resourcefulness, and determination. Pasteur performed more than one service for his fellow man. In his own words, "But of this we may be sure,—that science in obeying the laws of humanity, will always labor to enlarge the frontiers of life." In studying their lives, there appears to us the discriminating insight and practical thoroughness of Lavoisier, the broad grasp of Mendelejeff, the patience and penetration of Madame Curie, and the brilliant technique of Ramsay, Rutherford, and Richards in the search for truth. There are still scores of others whose names might be mentioned—men whom we can really think of as being the pioneers of to-day's civilization.

Chemistry is so closely connected with many phases of modern life that almost every one should have some knowledge of it. Its application to the problems of industry and recently to those of warfare has attracted unusual attention and interest. It plays no less important a part in the sciences of agriculture and medicine. For these reasons a knowledge of the *fundamental facts of chemistry* and an ability to use them should form a part of the equipment of our youth.

It is a most interesting fact that the chemist in his laboratory is all the time teaching processes which are used in great industrial plants where millions of dollars are expended and thousands of men employed. Some of the most common chemical materials, such as sulphuric acid, soda, and steel, are literally foundation stones of modern civilization.

Chemistry is one of the most important of the sciences, and during our study of it this year, we shall endeavor to learn order and system. We shall familiarize ourselves with certain fundamental laws and principles, and how they apply in our everyday life; we shall make our own observations of phenomena and reactions which will prove to us the facts and theories of our classroom study—and we shall attempt to organize and classify our knowledge so that it will help us in acquiring other knowledge.

#### QUESTIONS FOR STUDY

1. What do the chemists of large modern cities do to safeguard drinking water? What methods and materials do they use? What do you know about the supply of drinking matter in your native locality? Is it safe?
2. Has the science of chemistry done anything for the advancement of China these last eighteen years?
3. According to your present knowledge, what do you think is the most important contribution of chemistry to the world?
4. Why is chemistry of such great value in agriculture?
5. In what ways did chemistry play an important part in the World War?
6. Name five examples of chemical change (not mentioned in your textbook) with which you are familiar.

#### NOTES ON METHOD OF STUDY

1. Learn to make the best use of your textbook. Give attention to chapter and paragraph headings, and to statements which are

printed in italics. Use the index, give attention to footnotes, if any, and study out the meaning of pictures. How to use a book is an art that you should learn.

2. Study by yourself. Solve your own problems and do your own thinking; for in this way only will you grow as a scholar.

3. Studying is simply intensified reading. Therefore learn to read in such a way as to get the real thought. "One hour's bright, wide-awake, concentrated, interested study is worth a day's plodding."

4. Find and state the problems in the subject you are studying. You are well on the way to knowledge when you know exactly what you want to know.

5. Never ask another person to answer a question if by thinking you can answer it for yourself; but never go without a necessary bit of knowledge that you can gain by asking a question.

6. Remember that the book is only a guide to show you where and how to look for truth; and remember also that the truth is hidden away, not in the book, but in the object of your study. What that truth is, you must seek for yourself. It would be unfortunate for you if the book or your teacher were to give you information which by observation or experiment you can get for yourself.

#### LABORATORY NOTES

1. Follow laboratory directions closely, but not blindly.

2. Understand what an experiment is intended to teach or illustrate before attempting to perform it.

3. Notes on an experiment should be taken while performing an experiment or immediately after.

4. Observe carefully and take note of all that takes place during an experiment. *Think* about what you *see*.

5. The *quality* of work is far more important than the *quantity*.

6. Make your notes clear, concise, and neat.

7. Classify your notes under three separate columns opposite each other, thus:

OPERATION	OBSERVATION	CONCLUSIONS
Tell simply in your own words what you did.  Summary.	Tell simply what you saw happen.	Thing proved or illustrated, explanation, etc.

8. Keep these three columns strictly separate.

9. At the end of each experiment summarize briefly what that particular experiment was intended to prove.

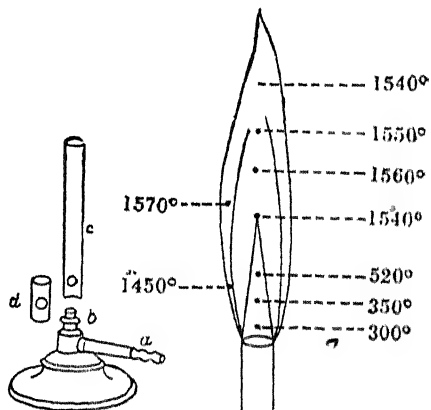
### CARE AND USE OF APPARATUS

**The Burner.** The Bunsen burner should burn with a clear, blue flame. The ordinary gas flame deposits soot on objects which it touches. The character of the flame may be changed by adjusting the quantity of air that enters the holes at the base of the burner.

**Heating Glassware.** Test tubes may be put directly in the flame; beakers and flasks should be protected by a wire gauze or asbestos mat. When glass apparatus contains a liquid, the flame should never extend above the liquid in the vessel. Never attempt to heat articles made of thick glass, such as bottles and battery jars, because the poor conductivity of glass causes unequal expansion and breakage.

**Heating Porcelain.** Evaporating dishes and crucibles can be heated to very high temperatures. Crucibles can be put directly in the flame, but evaporating dishes should be placed on wire gauze with asbestos center. In both cases the heat should be applied slowly at first.

**Setting Up Apparatus.** *a.* Have everything firmly arranged and securely placed.



From Black's "Laboratory Experiments"

Fig. 1. Bunsen Burner and Flame.

Figures indicate probable temperatures in degrees centigrade



*e.* See that rubber stoppers fit securely, but use care in pressing them into the necks of thin glass articles.

*f.* Never try to push a glass tube through the hole in a stopper. Moisten the end of the tube, and work it slowly through the hole, with constant turning.

*g.* The lower end of a thistle tube should dip under the surface of the liquid in the bottle or flask.

*h.* The bends in glass tubes should be rounding, not angular. The latter are likely to break, and the flow of a gas in them is partly obstructed.

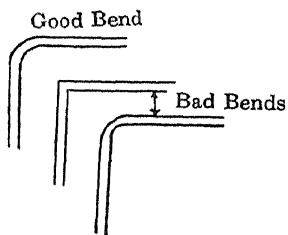


Fig. 3. Good and Bad Bends of Glass Tubing

*i.* Pay a good deal of attention to the "appearance" of your apparatus. Have vertical lines vertical, and horizontal lines horizontal.

*j.* After pouring from a reagent bottle, remove the drop of liquid usually sticking to its lip by touching it to the top of the receiving vessel.

## CHEMICAL CHANGE

### Experiment 1. Direct Combination of Two Elements.

**Apparatus.** Watch glass.

**Materials.** Mercury; powdered sulphur.

*a.* Place a small drop of mercury in a watch glass. Note the color and general appearance of the mercury.

*b.* Add a small pinch of powdered sulphur. Note the color and appearance of the sulphur before adding to the mercury.

*c.* Rub together, on the palm of the left hand with the fingers of the right hand, the mercury and the sulphur.

Describe what change has taken place, and the appearance and properties of the final product.

Summarize briefly the object of this experiment.

**Experiment 2. Decomposition of a Compound Formed by Heating a Metal in Air.**

**Apparatus.** Ring stand and clamp; hard glass test tube; rubber stopper; delivery tube; sink or deep vessel of water; test tube; splinter; Bunsen burner or alcohol lamp.

**Material.** Mercuric oxide.

a. Put about two grams of the red powder in a hard glass test tube fitted with a stopper carrying a delivery tube. Place the end of the delivery tube under the mouth of the test tube which is filled with water and inverted in the sink which is almost filled with water.

b. Heat *gradually* the hard glass test tube keeping flame in motion. Note any change in the color of the red powder. Then heat strongly until the red powder all disappears. When the water in the test tube has all been displaced, remove and test as in c. Collect a second tube of gas and test again. *Remove the delivery tube from the water before removing the flame. Why?*

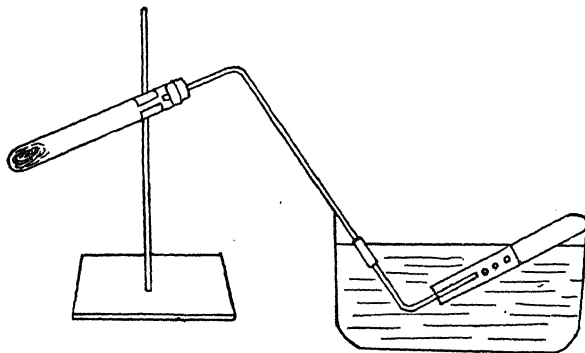


Fig. 4. Collection of Oxygen by Displacement of Water

c. Remove the filled test tube from the water, invert it, and insert a glowing splinter.

*Result? What is this gas?*

d. Scrape together the substance which has collected on the cooler portion of the hard glass test tube.

*What is this substance?*

*Of what two elements is the red powder composed?*

Summarize very briefly the object of this experiment.



## CHAPTER II

### OXYGEN

In our last laboratory lesson, we found that we could decompose a compound and obtain oxygen. There are a number of other compounds which we may also decompose, and from them obtain oxygen, namely, potassium chlorate, barium peroxide, and manganese dioxide. The names of the last two compounds you will notice indicate very clearly the content of oxygen.

The substance which we will use in our laboratory experiment is potassium chlorate. This white crystalline solid is thirty-nine per cent oxygen, and all this oxygen can be made to leave the compound if we apply heat. We decompose the compound much more easily—at a much lower temperature—when we add manganese dioxide to our potassium chlorate. The manganese dioxide, in this case, undergoes no chemical change; it *helps* the reaction to take place more easily and more quickly; we call a substance like this a *catalytic agent*.

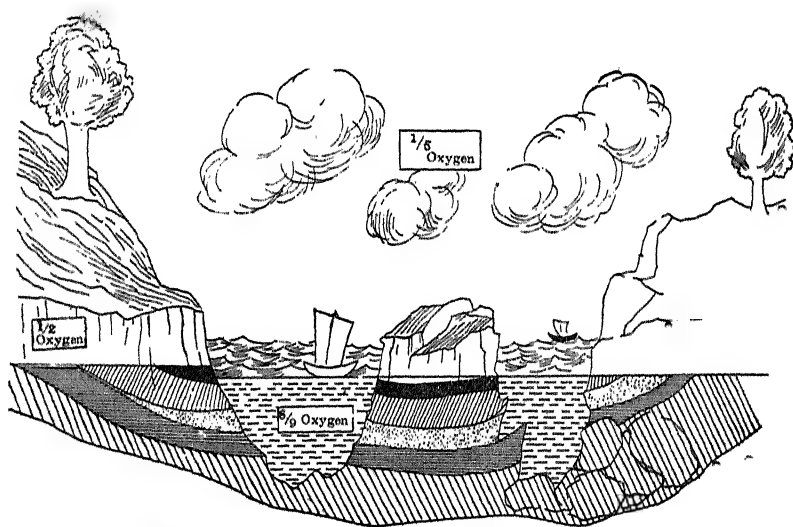


Fig. 5. By weight, oxygen makes up one half of the earth's crust, eight ninths of water, and about one fifth of the atmosphere

Substances like potassium chlorate, barium peroxide, and mercuric oxide, which give up some of their oxygen to other substances, or are sources of oxygen, are called *oxidizing agents*.

It is interesting to note that oxygen is the most abundant of all the elements, making up about *one half* of the solid part of the *earth*, about twenty-three per cent of the *air*, and eight ninths of the *water*, besides being a very necessary part of all living matter.

Oxygen unites very readily with most other elements. It can even be made to unite with nitrogen — an element which is one of the most inactive of the elements — making possible the production of many useful compounds.

The oxygen which we breathe into our lungs is carried to the different parts of the body by the blood stream. The waste products in the blood combine with this oxygen forming carbon dioxide, which is then carried back to our lungs and expelled when we exhale

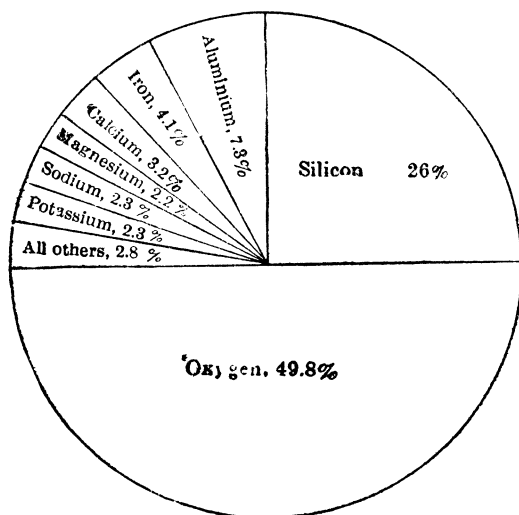


Fig. 6. Diagram Showing Approximately the Composition of the Earth's Crust, Including the Atmosphere

### QUESTIONS FOR STUDY

1. If all the oxygen in the world were suddenly taken away, would we be able to exist much longer? Why do you think as you do?
2. When a candle burns in the air should not new compounds be formed? What happens to them?
3. Why do we breathe more deeply in winter than in summer?
4. When a man swims under water, what does he do before he puts his head under the surface? Why does he do this?
5. Assuming the cost of potassium chlorate to be 50 cents per kilogram and that of mercuric oxide to be \$1.50 per kilogram, what is the cost of the weight of each required in the preparation of 10 grams of oxygen?

**Experiment 3. Preparation of Oxygen by the Decomposition of a Chlorate.**

**Apparatus.** Two test tubes; delivery tube; rubber stopper; ring stand with clamp; Bunsen burner; four wide-mouthed bottles; four glass plates for wide-mouthed bottles; sink; watch glass; funnel.

**Material.** Potassium chlorate; manganese dioxide; filter paper; splinter.

*In this experiment there is danger of the water "sucking back" into the hot test tube. Guard against this by removing the delivery tube from the water before the flow of gas stops.*

a. Mix thoroughly about 8 grams of potassium chlorate and 6 grams of manganese dioxide. Place in a test tube provided with a delivery tube. Clamp the test tube in a position convenient for heating. Carefully regulate the heating so as to cause a very gentle evolution of gas. *Do not heat the test tube sufficiently to make the flame yellow.*

b. Collect a test-tubeful of the gas and test with a glowing splinter. *Result?*

Collect the remainder of the gas in the 4 wide-mouthed bottles and keep for Experiment 4.

c. Remove the delivery tube and *allow the test tube to cool.* Nearly fill the test tube with hot water, close the mouth of the tube with the thumb, and shake the tube.

Pour the muddy liquid on a moistened filter paper fitted to a funnel. Collect the clear liquid which flows through the filter paper and funnel (called the filtrate) in a test tube.

Remove a small portion of the black residue from the filter, place it on a second watch glass, and set it aside to dry.

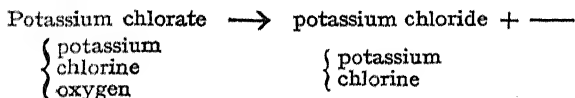
*Which of the original substances does the black residue resemble?*

Taste a crystal of potassium chlorate. Then taste the clear filtrate. Do they taste alike?

*Is the substance in the filtrate potassium chlorate or a different substance?*

*From which of the original substances was the oxygen probably derived?*

Complete the equation:



The manganese dioxide causes the oxygen to be liberated more regularly and at a lower temperature.

**Experiment 4. Formation of Oxides.**

**Apparatus.** Deflagrating spoon; Bunsen burner.

**Material.** Four bottles of oxygen; splinter of charcoal; magnesium ribbon; sulphur; red phosphorus.

a. Place a thin splinter of charcoal across the bowl of a clean deflagrating spoon. Heat the end of the splinter until it glows brightly, and immediately lower it into a bottle of oxygen.

*Does the charcoal burn with a flame?*

Compare the intensity of the action in oxygen with that in air.

*Of what elements does the gas formed probably consist?*

b. Twine a piece of magnesium ribbon around the rod of the deflagrating spoon, allowing the upper end to project slightly. Light the free end and lower the spoon into a bottle of oxygen.

Compare the combustion of the magnesium with that of the charcoal (carbon).

Compare the action in oxygen with that in air.

*What is the appearance of the oxide of magnesium? Does it look like the magnesium ribbon?*

c. Clean the spoon and place on it a small piece of sulphur. Heat the sulphur by directing the flame of the burner against it till it lights, and then lower it into the bottle of oxygen.

*Describe the burning of the sulphur in oxygen.*

When the mist (principally unburned sulphur) has disappeared, very cautiously smell the contents of the bottle.

*Name this gas.*

d. Clean the spoon and place on it a small pinch of red phosphorus. Heat the phosphorus till it lights, and put it into a bottle of oxygen.

*Describe the burning of phosphorus in oxygen.*

*Name this product.*

Make a general statement as to the relative rapidity of the burning of a substance in oxygen and its burning in air.

Give a general name for the product formed by burning an element in oxygen.

*Complete the equations :*

Carbon + oxygen  $\rightarrow$  \_\_\_\_\_ .

Magnesium + oxygen  $\rightarrow$  \_\_\_\_\_ .

Sulphur + oxygen  $\rightarrow$  \_\_\_\_\_ .

Phosphorus + oxygen  $\rightarrow$  \_\_\_\_\_ .

## CHAPTER III

### HYDROGEN

The last element we studied, i. e., oxygen, is found to be the most abundant of all the elements on the earth, but hydrogen is perhaps the most widely distributed of all the elements. There is very little of it free on the earth as an element, but it is exceedingly abundant as a constituent of many compounds. One ninth of the weight of water is hydrogen—and there is a large amount of water on this earth. All plant and animal tissues contain hydrogen; petroleum and its many related compounds contain hydrogen; many rocks contain hydrogen, and in many salt deposits, hydrogen is found. It exists in great quantities in the uncombined form in the atmosphere of the sun, in the stars, and throughout the universe; therefore, one can very well understand the importance of this element hydrogen.

The most common laboratory method for the preparation of hydrogen we will do in the laboratory, i. e., its formation by the reaction between an acid and a metal, in this case, zinc and hydrochloric acid. Acids, like water, contain hydrogen, and give it up readily when treated with certain metals under proper conditions. It is really the hydrogen which gives to these compounds called “acids” their characteristic properties.

When hydrogen is first liberated by the action of a metal on an acid, it has very different properties from those which it possesses after it has once been formed. While hydrogen gas, as we ordinarily know it, must be heated to a higher temperature before it will reduce the oxides of most metals, hydrogen which is just being formed will reduce many such substances, even at ordinary temperatures. In this condition, it is known as “nascent” hydrogen, “nascent” meaning “just being born.”

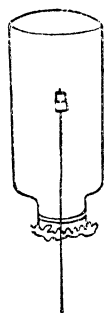


Fig. 7. Hydrogen burns at the mouth of the bottle, but does not support the combustion of the lighted taper in it.



Hydrogen has a great tendency to combine with oxygen, not only when oxygen is free as the element oxygen, but even when it is combined with other elements. Hydrogen is called a "reducing agent," i. e., it takes oxygen away from other substances. The process is called "reduction."

Hydrogen is the lightest of all substances known. It is nearly sixteen times lighter than oxygen, and oxygen, you remember, is only slightly heavier than air.

In our study of chemistry, hydrogen is taken as the unit of measurement of other substances, i. e., the standard by which we measure all other substances.

Whereas oxygen supported combustion but did not burn, in hydrogen we have just the opposite case, as it is a gas which burns with a hot, almost invisible flame—but it does not support combustion. The oxyhydrogen blowpipe, which is so constructed as to allow the hydrogen to burn in an atmosphere of oxygen, gives a very hot and intense flame, but now the oxyacetylene flame and the electric arc are beginning to take its place. The most important of the uses of hydrogen now are—to fill balloons and dirigibles, to harden oils, and in the manufacture of gaseous fuels.

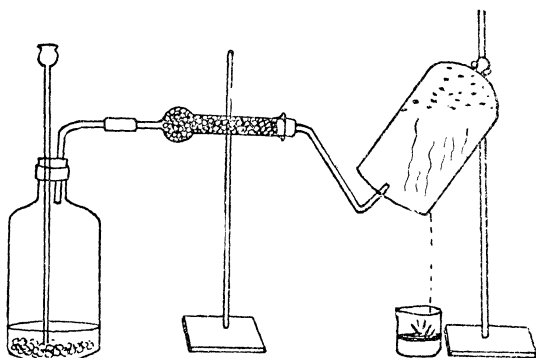


Fig. 9. Water is formed by burning hydrogen.

#### QUESTIONS FOR STUDY

1. Do we ever have oxidation without reduction?
2. In what properties are oxygen and hydrogen alike? In what are they unlike?
3. What property has hydrogen which makes it a dangerous substance with which to inflate balloons?
4. Could pure hydrogen be used in a Bunsen burner instead of illuminating gas?
5. How many grams of hydrogen can be made from 100 grams of water?



## LABORATORY REMINDERS

1. When you finish your laboratory work, turn off the gas, disconnect the apparatus, and leave your desk in good order. Leave all things which belong to the *laboratory* on *top* of the desks.

2. Keep your glassware clean and your locker in good order. There will be an inspection of lockers each month by the teacher in charge and those with disorderly lockers will be discredited accordingly.

3. If there is anything you do not understand do not hesitate to ask the teacher. He is always ready and willing to help if there is anything troubling you, provided you cannot solve the problem yourself.

4. Be very careful, in working with sulphuric acid, not to get any on your clothes or to spatter it on yourself or your neighbor. It is a powerful dehydrating agent and will make very bad burns, especially if the acid is concentrated.

**Experiment 5. Preparation of Hydrogen by the Reaction Between an Acid and a Metal.**

**Apparatus.** Flask for generating hydrogen; 2-holed stopper; thistle tube; delivery tube; sink; 3 wide-mouthed bottles; 3 glass plates; watch glass; beaker; ring stand with ring; Bunsen burner.

**Material.** Zinc; dilute sulphuric acid; copper sulphate solution.

**Caution!** A mixture of hydrogen and oxygen (or air) is dangerously explosive. Have no light or burner near your generator. Collect the gas in test tubes until you find that the gas burns quietly when a flame is applied to it.

*a.* Use a flask provided with a two-holed stopper carrying a thistle tube and a delivery tube for a generator.

*Why must the end of the thistle tube dip below the surface of the liquid?*

Have 3 bottles filled with water, standing inverted in the sink. Put about 20 grams of granulated zinc into the generator, and pour through the thistle tube dilute sulphuric acid until one fourth of the bottle is filled.

*Result?*

If the action is slow in starting, add a few drops of copper sulphate solution through the thistle tube.

*Observe and describe the action in the generator.*

*Under no circumstances add more acid, nor in any way interfere with the generator without consulting the instructor.*

b. Collect the gas in a test tube. As soon as the test tube is filled, hold its mouth downward to a small flame. Continue to collect and test the gas in this manner until a portion burns quietly. The hydrogen is now ready to be collected for the next experiment.

Fill 3 bottles with the gas. Leave them standing in the sink, or cover them with glass

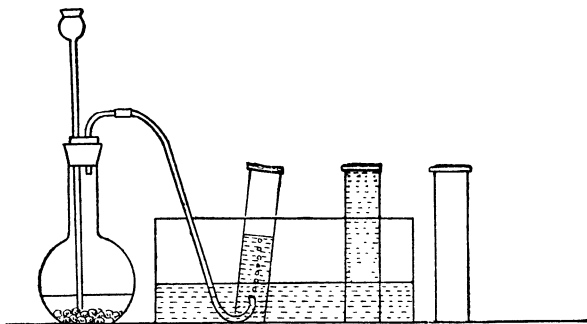


Fig. 10. Preparation of Hydrogen by the Action of Metals on Acids

plates and set them mouth downward on the desk.

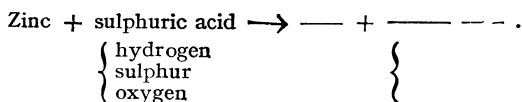
Proceed to the next experiment, finish it, and come back and do part c.

c. Filter a few drops of the liquid in the generating bottle into a watch glass. Place this *on top of a beaker* one third full of water and boil the water until a solid has appeared on the watch glass.

*Examine this solid and describe its appearance.*

The compound is zinc sulphate, which is composed of zinc, sulphur, and oxygen.

Complete the equation :



*From what material does the hydrogen come?*

*Why is the action called a replacement action?*

*Why was the gas collected in test tubes and repeatedly tested in the first part of the experiment?*

### Experiment 6. Properties of Hydrogen.

**Material.** Wooden splinters.

a. Holding a bottle of hydrogen mouth downward, thrust into it a lighted and burning splinter.

*What happens to the flame of the splinter?*

*What is occurring at the mouth of the bottle?*

Slowly withdraw the taper from the bottle.

*Explain the result.*

b. Lift a bottle of hydrogen, and hold it, mouth downward and uncovered, for a full minute by the watch. Then hold the mouth of the bottle to a flame.

*Result?*

Hold another bottle, uncovered and mouth upward, for a full minute, and again hold the mouth to a flame.

*Result?*

*Is hydrogen heavier or lighter than air? Give reasons for your answer.*

Now return to part c of the last experiment.

## CHAPTER IV

### COMPOSITION OF WATER

So far we have studied and worked with two *elements*, hydrogen and oxygen. The best-known chemical *compound* is water.

Water, in the free condition, covers three fourths of the surface of the earth. It exists in large quantities in the atmosphere in the form of water vapor. It exists in combination with a large number of substances as water of crystallization.

TABLE OF WATER CONTENT OF COMMON ANIMALS

<i>Animal</i>	<i>Water Content</i>
Fat Ox . . . . .	45.5 per cent
Fat Calf . . . . .	63 per cent
Fat Pig . . . . .	41 per cent
Lean Pig . . . . .	55 per cent
Fat Sheep . . . . .	43 per cent
Lean Sheep . . . . .	51 per cent
Chicken Flesh . . . . .	74 per cent
Goose Flesh . . . . .	42 per cent
Turkey Flesh . . . . .	55 per cent
Brook Trout . . . . .	78 per cent
White Fish . . . . .	70 per cent
Lobster (Exclusive of Shell) . . . . .	79 per cent
Oyster (Exclusive of Shell) . . . . .	81 per cent

TABLE OF WATER CONTENT OF VARIOUS PARTS OF AN OX

<i>Part</i>	<i>Water Content</i>
Brain . . . . .	80 per cent
Heart . . . . .	63 per cent
Muscle of the Loin . . . . .	62 per cent
Kidney . . . . .	76 per cent
Liver . . . . .	71 per cent
Sweetbreads . . . . .	71 per cent
Lungs . . . . .	80 per cent

It forms an essential part of all living matter. More than two thirds of the human body is water, and the animal and vegetable food which we eat contains scarcely less water in proportion to solid matter. We can thus see why animals can live without food much longer than without water, and why water is absolutely essential to vegetable and animal life.

All natural waters contain impurities; not only that which is thrown into water artificially, as by the drainage of human habitations, but also the matter which is dissolved from the rocks and soil by the waters on their way to larger bodies of water or to the sea. Rain water is the purest of natural waters, but even it may be contaminated by the impurities in the air, such as particles of dust, dirt, solid matter or gases during its passage through the atmosphere.

We have seen that oxygen and hydrogen can be obtained from water, but this does not show that water contains only these two

elements. To answer this question two general methods are available. First, decompose water, and see if anything but hydrogen and oxygen are obtained. Second, cause oxygen and hydrogen to combine and see whether water is formed.

The most convenient means of decomposing water is the electric current. The process is known as electrolysis. The only products obtained by the electrolysis of water are the two gases, oxygen and hydrogen, oxygen being set free at the anode, and hydrogen at the cathode. That these are oxygen and hydrogen, respectively, can

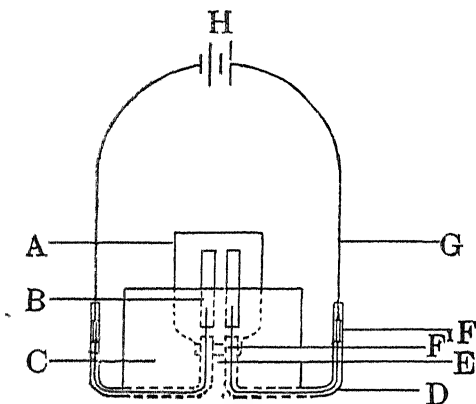


Fig. 11. Simple Apparatus to Illustrate Electrolysis of Water

A. Glass bottle with bottom off, containing slightly acidulated water.

B. Test tubes for collection of hydrogen and oxygen.

C. Wooden block into which bottle is set.

D. Glass tubing containing mercury.

E. Hard rubber cork.

F, F'. Wax.

G. Copper wire.

H. 2 dry cells.

be shown by the fact that the former will ignite a match which has just been extinguished, and the latter will burn with the characteristic

hydrogen flame. If we wish to know the volumes of these gases set free from water, we can collect and measure them—and we shall find that the volume of the hydrogen is always twice that of the oxygen. This breaking up of a compound to find its constituent elements is called *analysis*.

Whenever two volumes of hydrogen are mixed with one volume of oxygen and the gases made to combine by means of an electric spark, or by rise in temperature, all the hydrogen and all the oxygen are used up and water is formed. If more than two volumes of hydrogen are used, all

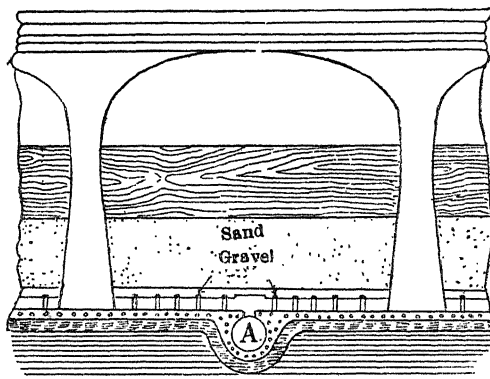


Fig. 12. A Covered Sand-Filter Bed

After filtering through the sand and gravel, the purified water passes into the porous pipe A, from which it is pumped into the city mains.

the oxygen will be used up, and the excess of hydrogen will remain uncombined. If less than two volumes of hydrogen are used, all the hydrogen will be used up and the excess of oxygen will remain. The combination of the *elements* hydrogen and oxygen to form the *compound* water is called the *synthesis* of water.

Water may be purified by *filtration* and *distillation*. Water which has been *filtered* has had the solid impurities caught and retained. In the laboratory we use filter paper; in the household filters the suspended matter is caught by means of a porous cylinder,—and on a large scale, as in the city

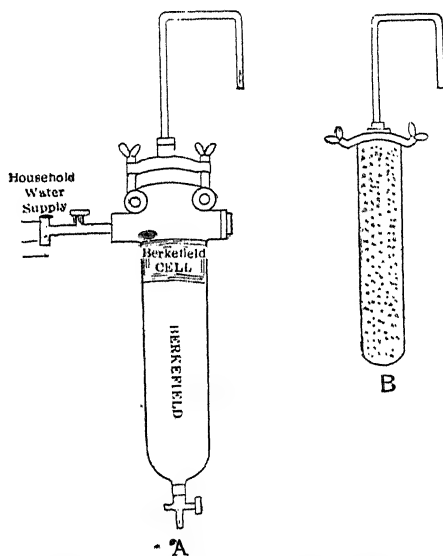


Fig. 13. Ordinary Household Filter

A. Outside view.

B. Porous cup inside, through which water is forced, to be purified.

waterworks, the water is allowed to pass slowly through beds of gravel and sand. The soluble impurities, however, still remain in the water.

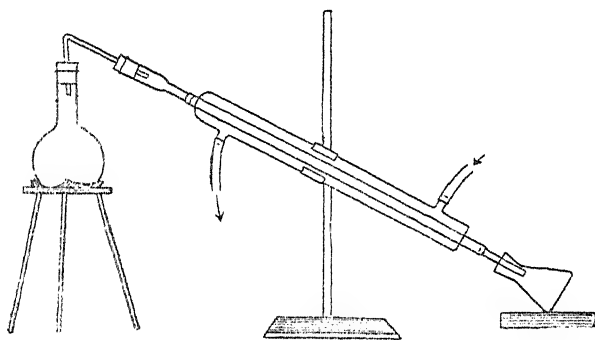


Fig. 14. Liebig Condenser

Arranged for the distillation of water or other liquids.

volatile will remain behind. In distillation all the dissolved and suspended material can be removed.

In the process of *distillation*, a liquid is changed into a vapor and the vapor is then condensed to the liquid state. When water is distilled, all impurities more volatile than the water will appear in the first portion of the *distillate*; all much less

### QUESTIONS FOR STUDY

1. Why are natural waters not chemically pure? What impurities are then in rain water? in well water?
2. How may it be shown that water is a compound of hydrogen and oxygen?
3. Would ice frozen from impure water necessarily be free from disease germs?
4. How could you determine whether a given sample of water is distilled water?

### Experiment 7. Distillation of Water.

**Apparatus.** Flask 250 c. c.; 2 glass bends; wire gauze; clamp; thistle tube; 2-holed stopper; beaker; test tube; Bunsen burner; watch glass; dropper.

**Materials.** Copper sulphate crystals; solution of phenolphthalein; ammonia water (concentrated).

Insert thistle tube and short glass bend in stopper which fits into the flask, and then connect with a short rubber tube to the long glass bend which leads into the test tube immersed in the beaker of cold water.

- a. Put into the distilling flask a few crystals of copper sulphate and add 100 c. c. of water. Heat the contents of the flask to boiling.

*Does the copper sulphate dissolve? Give a reason for your answer.*

Let the distillate (the condensed steam) collect in the test tube immersed in the beaker of cold water.

*Does it contain any copper sulphate? Give a reason for your answer.*

Dissolved solids in general act in distillation as the copper sulphate does.

A substance which readily changes into a gas on being heated is said to be volatile.

*Is copper sulphate volatile under the conditions of this experiment? How do you know? Can water be purified from nonvolatile impurities by distillation?*

b. Add a drop of phenolphthalein solution to some pure water with a dropper.

*Result?*

This test is used to indicate the presence of ammonia in water in this experiment.

Prepare another mixture of ONE DROP of concentrated ammonia water in 200 c. c. of pure water. Pour it into a clean flask and distill. Collect a little of the distillate in a clean test tube. Add a drop of phenolphthalein solution.

*Does the distillate contain ammonia? Is ammonia volatile?*

Continue the distillation for 2 or 3 minutes and then collect a second portion of the distillate.

*Does it contain any ammonia?*

*Does the first portion of the distillate contain more or less ammonia than the second portion? Can water be readily purified by distillation from a volatile impurity like ammonia?*

c. Determine, by evaporation on a watch glass over steam, whether water from the faucet contains nonvolatile impurities.

*Result?*

Suggest a method by which salt water can be converted into fresh water.

Summarize very briefly the object of this experiment.

Make a drawing of your apparatus.



## CHAPTER V

### WATER AND SOLUTION

*Solution*, or dissolving, takes place when substances are mixed in such a way that *the matter of each is distributed uniformly through that of the other*. Examples of common solvents are *water*, *alcohol*, and *ether*.

Water has a remarkable power to dissolve other substances which are brought in contact with it. Indeed, of all known substances it is the best solvent. The importance of solution for chemistry cannot be overestimated. This becomes obvious when we consider that *most chemical reactions take place in solution*. Comparatively few solid substances are capable of reacting with other substances in the solid state. Were it not for solution the whole science of chemistry would be very different from what it is to-day, and far less interesting. Three fourths, and probably a much larger proportion of the chemical reactions with which we are now familiar would not take place at all.

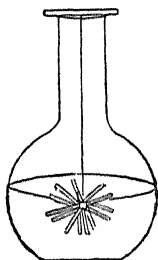


Fig. 15. The addition of a single crystal to a supersaturated solution will cause rapid crystallization.

Water dissolves to a greater or less extent not only most solid substances which are brought in contact with it, but also most liquids and gases.

When water has dissolved a certain amount of a given substance, but is still capable of taking up more of it, the solution is said to be *unsaturated*. When water has dissolved all of a given substance which, at the given temperature it can take into solution, the solution is said to be *saturated*. When water contains more of a substance than it can hold in a stable condition, the solution is said to be *supersaturated*.

The solubility of a substance depends not only upon the solvent used, but also upon the *temperature*. As a rule, solids are more soluble in hot than in cold liquids, while the *reverse* is true of *gases*.

For any gas, the solubility varies with the pressure,—the greater the pressure the greater the solubility.

The *freezing point* of water is lowered by dissolved substances, and this is perfectly general no matter what the substance which is dissolved in the water. Similarly, the *boiling point* of water is raised by the presence of substances dissolved in it.

## QUESTIONS FOR STUDY

1. Why do underground waters contain more gaseous substances in solution than waters on the surface of the ground?
2. Why is it that an insoluble substance like sand has no taste?
3. How could you separate a mixture of white sand and common salt so as to recover all of each in a dry condition?
4. How could the presence of air dissolved in water be detected?

**Experiment 8. Solution and Suspension.**

**Apparatus.** Beaker 250 c. c.; large beaker; two test tubes; stirring rod.

**Material.** Rock salt or alum; fine salt; calcite; powdered sodium chromate; precipitated chalk.

*a. Soluble and insoluble substances.*

Put a piece of rock salt about the size of a pea into a test tube of water. In another test tube of water put a lump of calcite of equal size. Close the mouth of each of the test tubes and shake them.

*Which substance disappears in the water?*

*Which substance has not dissolved?*

*How can you tell this from the shape of the pieces of solid remaining in the water?*

*Name the soluble substance.*

*b. Distribution of the solute in the solvent.*

Put a half teaspoonful of table salt in a beaker of water and stir the mixture.

*Is the solution clear and transparent?*

Dissolve a half teaspoonful of powdered sodium chromate in a beaker of water by stirring thoroughly.

*Is the solution clear?*

*Is the solution transparent?*

*What does the sameness of color throughout the solution show about the distribution of the dissolved substance (the solute) in the water (the solvent)?*

Allow the solution to stand for a short time.

*Does the sodium chromate settle to the bottom of the beaker?*

c. *Suspension.*

In a beaker of water put a teaspoonful of precipitated chalk. Stir thoroughly.

Such a mixture is called a *suspension*.

Allow the jar to stand for a time.

*What does the solid in suspension tend to do?*

*What difference do you notice between a solution and a suspension?*

State three characteristics of a true solution that have been observed in this experiment.

## CHAPTER VI

### THE LAW OF MULTIPLE PROPORTIONS, AND WATER OF CRYSTALLIZATION

*Chemical laws* are general statements of facts learned from many experiments. They are statements of fundamental principles to help us in our study of chemical changes. They are generalizations which repeated observations have shown to be true, and they have been of great importance in the development of the science.

We found that water was composed of 1 part of hydrogen to 8 parts of oxygen, by weight, or, if measured in volumes, two volumes of hydrogen to one volume of oxygen; and that the proportions of hydrogen and oxygen never varied in the compound water, no matter from what part of the world we took our sample. The compound water illustrates for us very well the Law of Definite Proportions which states that *a given chemical compound, no matter what its source, will always contain the same elements united in the same proportions*.

But we find another substance, hydrogen peroxide, which is made up also of hydrogen and oxygen. The proportions, though, are 1 part of hydrogen by weight, to 16 parts of oxygen by weight, whereas in water, you remember, it was 1 part by weight of hydrogen to 8 parts by weight of oxygen. The weights of hydrogen are the same in both compounds, but you notice that there is just twice as much oxygen by weight in the peroxide as in the water.

These two compounds illustrate for us the generalization which is called "The Law of Multiple Proportions" which states that *"if two elements form several compounds with each other, the different weights of one element which combine with a fixed weight of the other bear a simple ratio to one another."* This law has been confirmed by the work of the past century and is one of the fundamental principles of chemistry.

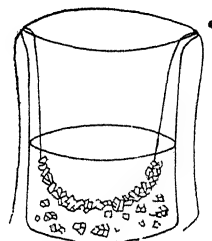


Fig. 16. Crystals of Alum Crystallizing on a Thread Suspended in a Concentrated Solution of Alum

By *water of crystallization*, we mean the water with which some substances combine when they crystallize from water solution. When substances containing water of crystallization are heated, they usually *melt* while the water escapes, and then assume the solid form again.

The amount of crystal water which will combine with a given weight of the *anhydrous* (i. e., water-free) substance is definite for each substance. The loss of crystal water by a substance is accompanied by a loss of crystalline structure and by other changes in properties.

### QUESTIONS FOR STUDY

1. How many grams of water are formed by the combustion of 10 grams of hydrogen in air?
2. How does the composition of hydrogen peroxide illustrate the Law of Multiple Proportions?
3. Name three crystalline substances.

### Experiment 9. Water of Crystallization.

**Apparatus.** Horn pan balance; 2 watch glasses; 8 test tubes; test-tube rack; Bunsen burner.

**Materials.** Crystallized sodium sulphate; copper sulphate; potassium chlorate; zinc sulphate; barium chloride; potassium sulphate; potassium nitrate; alum.

a. Place a small watch glass on the pan of a balance and then add enough crystallized sodium sulphate to balance the weight in the other pan. Set the crystallized sodium sulphate aside for half an hour.

*Note and record any changes in appearance.*

Again place the watch glass and contents on the balance.

*What does the change in weight indicate?*

b. Heat a few crystals of sodium sulphate carefully in a dry test tube, *holding the tube in a horizontal position*.

*What collects on the walls of the test tube? To what is the change in weight noticed in (a) probably due?*

c. In another test tube, gently heat a crystal of blue vitriol (copper sulphate) over a small flame until a white substance is produced.

*What important changes have taken place in the color, form, and composition of the crystal?*

After cooling, dissolve the residue in the bottom of the tube in a *few drops* of hot water. Pour the solution on a watch glass and allow it to stand. Compare the final product with the original substance.

*Result?*

*What took place when the copper sulphate was heated?*

*What happened when the residue was dissolved and allowed to cool?*

d. In separate *dry* test tubes gently warm a few crystals of potassium chlorate, zinc sulphate, barium chloride, potassium sulphate, potassium nitrate, alum.

Record the results in the following tabular form:

SUBSTANCE HEATED	IS WATER DEPOSITED IN COOL PORTION OF TUBE?	APPEARANCE OF RESIDUE

## CHAPTER VII

### ATOMS AND MOLECULES

*Molecules* and *atoms* are *not identical*; for, while atoms are thought of as the very smallest particles into which matter is capable of being divided, molecules are held to be *aggregations of atoms* which form the physical *units* of matter. The atoms composing a molecule do not (usually) part company when matter undergoes a *physical* change

The molecules of *elements* consist of atoms of only *one* kind, while the molecules of *compound substances* contain atoms of *two* or *more* kinds.

The view that matter is composed of indivisible particles or atoms which have definite weights, and that *chemical action* takes place between these particles, was to Dalton, the formulator of the Atomic Theory, the only rational explanation of the laws of definite and multiple proportions and the law of combining weights. If water is composed of such indivisible particles or atoms, then a definite number of atoms of one substance combines with one atom of another substance to form a definite *molecule* of the compound, and we have the law of definite proportions. *One atom* of *one* substance may combine with *one atom* of *another* substance, or a *number of atoms* of *one* substance may combine with *one of another* to form a molecule, but the number must be a simple, rational, whole number; whence the Law of Multiple Proportions.

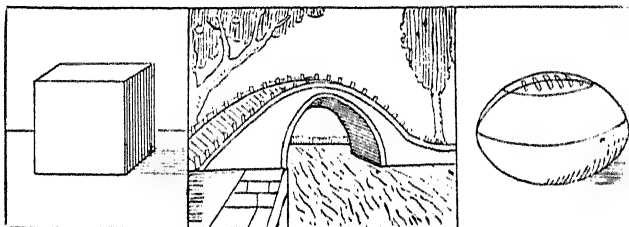


Fig. 17. In a solid the molecules are confined in their places; in a liquid they hold together but move easily over each other; in a gas they fly apart.

Since the atoms have definite weights, and the laws of definite and multiple proportions are true, the law of combining weights follows as a necessary consequence of the atomic theory.

The question as to the *size or mass of an atom* is one which is still open to some doubt. We know that they are inconceivably small. This is shown by the fact that certain substances will continue to give off odors for a long time, which fill a large space, and still not lose appreciably in weight. The odoriferous particles must be present in every part of the space, and although the substance will continue to fill this space with such particles for months or longer, the amount of matter which has volatilized is scarcely weighable. This shows the almost unlimited divisibility of which matter is capable; and by definition the atom is indivisible. The same fact is brought out by dissolving certain coloring matters such as aniline dye in water. Very small amounts of such substances can impart an appreciable color to comparatively enormous volumes of water. The coloring matter must be capable of almost unlimited divisibility in order that this may be effected.

Perhaps, on the whole, the best idea of the size of atoms has been furnished us by Lord Kelvin, of England. In his own words: "Imagine a raindrop or a globe of glass as large as a pea, to be magnified up to the size of the earth, each constituent being magnified in the same proportion. The magnified structure would be coarser grained than a heap of small shot, but probably less coarse grained than a heap of cricket balls."

#### QUESTIONS FOR STUDY

1. Of what importance is the Atomic Hypothesis to scientific progress?
2. How is the Law of Definite Proportions explained according to the Atomic Hypothesis? the explanation of the Law of Multiple Proportions? the Law of Combining Weights?

#### **Experiment 10. Temperature and Relative Solubility.**

**Apparatus.** 4 test tubes; Bunsen burner.

**Materials.** Powdered potassium nitrate; fine salt.



a. Fill a test tube with powdered potassium nitrate (saltpeter). Add the potassium nitrate, a very little at a time, to another test tube one fourth full of cold water, shaking the test tube after each addition. Continue to add the potassium nitrate until a very little of the solid remains after thorough shaking.

Make a drawing of your test tube (exact size, cross section) in your notebook, and mark with a pencil on this drawing, the upper level of the nitrate remaining in the supply tube.

Since the cold water has dissolved all that it can of the potassium nitrate, it is said to be *saturated* with respect to that substance at that existing temperature.

Heat in the Bunsen flame the cold saturated solution of potassium nitrate just prepared, and then add more nitrate from the supply tube. Shake the tube after each addition as before, and dissolve in the water as much as you can of the nitrate. During the process keep the liquid hot by heating occasionally in the flame, *but take care not to boil off any of the water*. As soon as the hot water is saturated with nitrate, set the solution aside to cool for later examination. Mark, in the diagram in your notebook, the upper level of the nitrate remaining in the supply tube, thus showing the amounts of the nitrate dissolved in equal volumes of hot and cold water. Label carefully on your diagram.

*Is the amount of potassium nitrate dissolved in the cold water greater or less than the amount dissolved in the same quantity of hot water?*

b. Examine the solution set aside to cool.

*Describe what has happened.*

*When the liquid comes to the temperature of the room, is it a saturated or an unsaturated solution?*

*How can you determine this?*

c. Repeat part a, using fine salt (sodium chloride) instead of potassium nitrate.

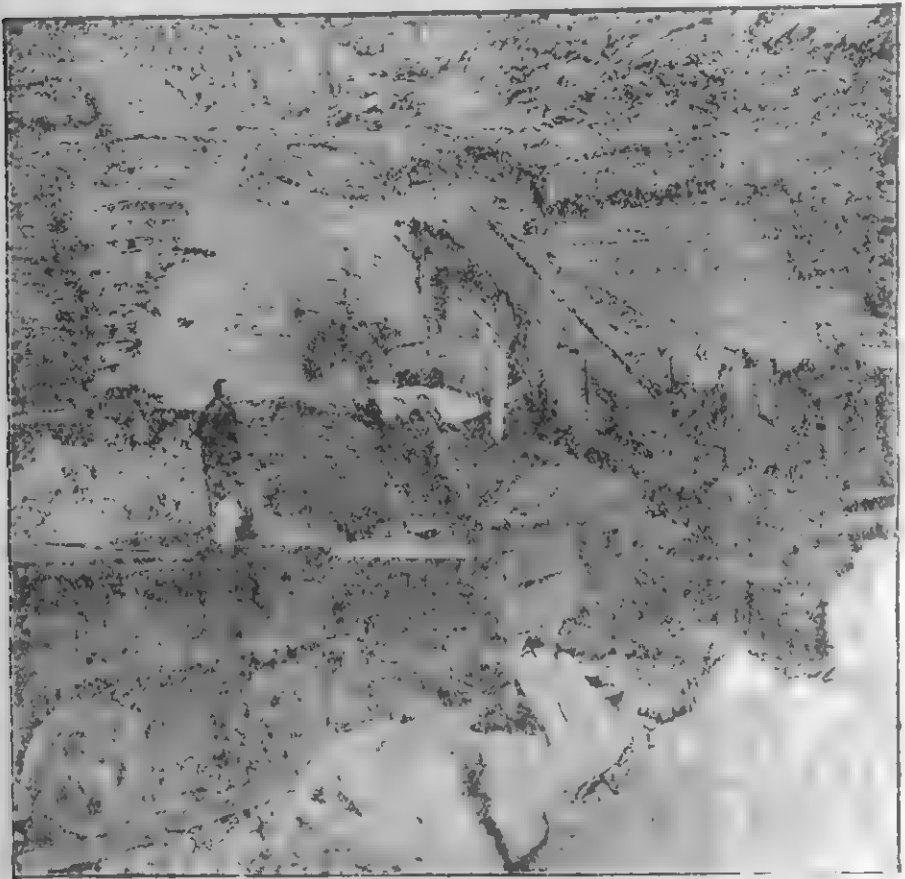
*Which is increased the more by increase in temperature, the solubility of sodium chloride or that of potassium nitrate?*

*Make a statement concerning the relative solubility of salt in hot and cold water.*

## CHAPTER VIII

### CHLORINE

Chlorine does not occur in the free condition in nature. This is due in part to its great chemical activity. If once set free it would quickly combine again with other substances. It occurs in combination with many other elements, such as magnesium, potassium, silver, lead, but



*The Detroit News*

Fig. 18. Mining Salt on the River Rouge, near Detroit, U. S. A.

Michigan owes her salt deposits to a great inland sea that occupied her present area some millions of years ago. Only after lumbering had ceased to be a bonanza did her people turn seriously to mining their saline wealth.

especially in combination with the element sodium as sodium chloride or common salt in our everyday use. The chlorides of all the above-mentioned elements are readily soluble in water, except the chlorides of lead and silver. The soluble chlorides cannot exist on the surface of the earth, where they are subjected to the influence of water, but pass into solution and are swept down to the sea. This accounts for the large amounts of chlorine in sea water, mainly in the form of potassium, magnesium, and sodium chlorides.

In certain protected localities, however, which are not readily accessible to water, as in the great salt beds of the earth, the chlorides may remain in solid form. As an example, take the great deposits at Stassfurt in Germany, and Salzburg. The deposits were made by the evaporation of the seas which once covered these regions, and which contained the various salts in solution.

Most of the chlorine used is now prepared by the electrolysis of concentrated solutions of potassium or sodium chloride.

The yellowish-green gas chlorine is, chemically, one of the most active substances known. It combines with nearly all the elements and with many compounds by simple contact, often with the evolution

of much heat and even light and in some cases almost with explosive violence. The best method of collecting chlorine for experimental purposes is by displacement of air. Being heavier than air the chlorine gas is conducted to the bottom of the vessel, and the latter is displaced upward by the heavier chlorine.

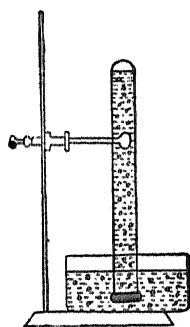
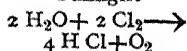


Fig. 19. The Action of Chlorine on Water in the Sunlight



Chlorine is readily soluble in water, the resulting solution being known as *chlorine water*. Chlorine water, if kept in the dark, is a stable substance, but if exposed to the light, a deep-seated change takes place. The chlorine acts chemically on the water, combining with the hydrogen (which it always has a great tendency to do) and liberating oxygen. The resulting solution contains the hydrochloric acid formed

while nascent oxygen is liberated. Such chemical reactions which are brought about by the action of light are known as *photochemical reactions*. Since oxygen is liberated, chlorine is known as a *strong oxidizing agent*.

Its oxidizing power renders chlorine one of the very best *bleaching agents* which is at our disposal. The oxygen which is set free when chlorine acts on moisture oxidizes organic coloring matter, and leaves behind the colorless substance. This can be illustrated by bringing into the presence of chlorine gas some *moist* flowers or a *moist* piece of colored cloth, when the color will disappear in a short time.

Chlorine is used in large quantities as a *bleaching* and *disinfecting agent* and is generally made for these purposes from *bleaching powder*, or "chloride of lime." Bleaching powder is a white substance formed by the action of chlorine upon "slaked" lime (calcium hydroxide), and is easily decomposed by acids, even by the carbon dioxide of the air, with evolution of chlorine. Fabrics to be bleached by the chlorine process are, therefore, immersed in a bath of dilute acid, and then in one of chloride of lime. In this way chlorine is set free in immediate contact with the coloring matter of the cloth, and bleaches it.

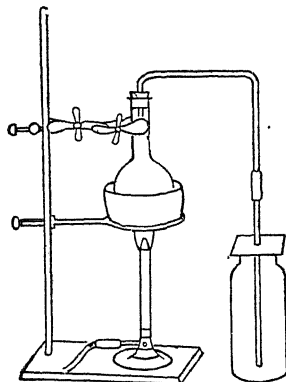


Fig. 20. Laboratory Preparation of Chlorine by the Reaction Between Salt and Sulphuric Acid, with Manganese Dioxide as a Catalytic Agent

### QUESTIONS FOR STUDY

1. What do we mean by a photochemical reaction?
2. What is the reaction of chlorine toward hydrogen combined with another element, as in water?
3. What is the action of the chlorine in bleaching powder?
4. In the preparation of chlorine, what advantages are there in treating manganese dioxide with a mixture of sodium chloride and sulphuric acid rather than with hydrochloric acid?

### Experiment 11. Bleaching Powder.

**Apparatus.** Watch glass; 2 small beakers; test tubes; dropper; funnel; filter paper; graduate.

**Materials.** Fresh bleaching powder; litmus solution; hydrochloric acid (dilute); sodium carbonate solution.

a. Note the smell of fresh bleaching powder. *What does it smell like?*

b. Dissolve 10 grams of it in 150 c. c. of water. Filter the resulting solution?

*What does the solution smell like? Why?*

c. To 5 c. c. of the solution on a watch glass, add a few drops of hydrochloric acid

*Judging from the odor, what has been liberated?*

Hypochlorous acid has been formed by the reaction between the hydrochloric acid and the bleaching powder solution. Bleaching powder is calcium hypochlorite, calcium hydroxide which has absorbed chlorine. Hypochlorous acid and its salts are very unstable, and in the presence of any material that will absorb oxygen, they readily liberate oxygen and break down into chlorides.

d. To 10 c. c. of the bleaching powder solution add a drop or two of litmus solution. Add a small piece of colored cloth to it.

*Result?*

Make a similar test with bleaching powder solution and indigo solution.

*Result?*

Make a similar test with bleaching powder solution and hydrochloric acid.

*Result?*

e. To 20 c. c. of bleaching powder solution add sodium carbonate solution till a precipitate no longer forms. Filter off the precipitated calcium carbonate. Test the resulting solution (often called Javelle water) for bleaching action on colored cloth.

Add a little hydrochloric acid.

*Result? Why?*

Summary.

## CHAPTER IX

### HYDROCHLORIC ACID

We have studied the relations by volume in which hydrogen and oxygen combine, and the ratio between the volumes of the gases which enter into combination and the volume of the product formed. It will be remembered that one volume of oxygen combines with two volumes of hydrogen, and forms two volumes of water vapor. The relations in the case of hydrogen and chlorine are even simpler. When one volume of hydrogen is mixed with one volume of chlorine and combination takes place, all of both gases are used up, and just two volumes of hydrogen chloride are formed. The law of the simple volume relations in which gases combine holds here even more strikingly than in the case of oxygen and hydrogen, there being no contraction in volume when the gases hydrogen and chlorine combine, and further, these gases combine in the simplest ratio by volume, i. e., equality.

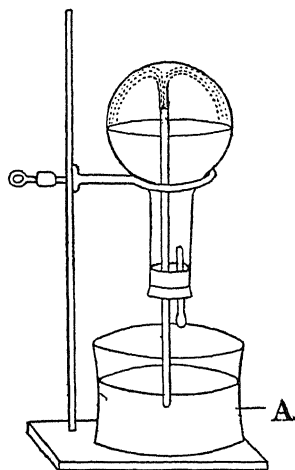
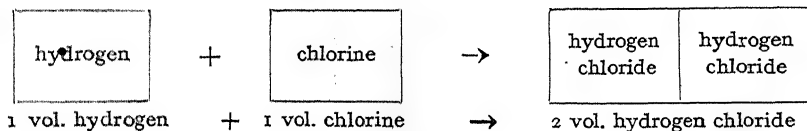


Fig. 21. The Hydrogen Chloride Fountain

Hydrogen chloride is very soluble in water. (If the liquid in A is colored blue with litmus, the experiment is much more striking.)



When the dry hydrogen chloride gas is dissolved in water, hydrochloric acid (muriatic acid) is formed. Dry hydrogen chloride gas possesses no characteristic acid properties.

Hydrochloric acid is typical of a large class of chemical compounds known as *acids*, all of which contain hydrogen. In a dilute water solution the hydrogen can be liberated by metals, the action being a case of replacement. Some substances, like sugar, contain hydrogen

which cannot be replaced by metals; they are therefore not acids. Water solutions of all acids have a peculiar sour taste, and all turn a certain natural dye, called litmus, from *blue to red*. Some of the commoner acids besides hydrochloric and sulphuric are nitric, acetic (which gives vinegar its sour taste), and citric acid (which is one of the acids present in lemons). Another characteristic of acids is the rapid reaction with certain compounds known as bases.

Hydrochloric acid gas is colorless and heavy, fumes in moist air, and dissolves readily in water. It is found only in small amounts in nature, e. g., in volcanic gases and in some springs. It makes up about 0.22 of 1 per cent of the gastric juice. This acid is produced in the body from the salt that is taken with food.

Concentrated hydrochloric acid is manufactured in considerable quantities for use in the industries and in laboratories. It is used by tinsmiths and plumbers for cleaning metal surfaces which are to be soldered.

The method used in manufacturing hydrogen chloride on a commercial scale is not different in principle from the method which we use in our laboratory. When sulphuric acid and sodium chloride (common salt) react, the hydrogen of the acid combines with the chlorine of the salt to form the gas hydrogen chloride, which is given off. There is left behind in the flask or retort a white solid which is called sodium sulphate.

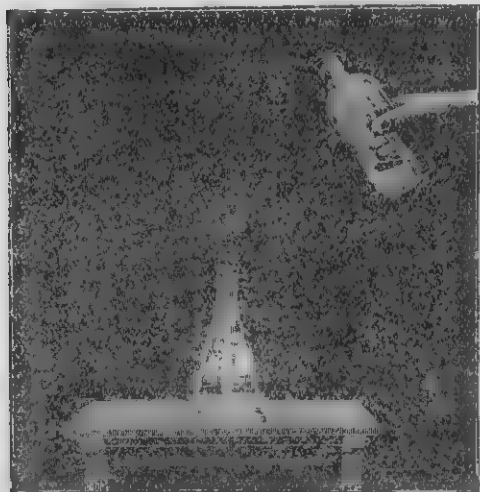


Fig. 22. Pouring silver nitrate into sodium chloride precipitates silver chloride.

A chloride is a compound of chlorine and some other element. All the common chlorides except three are soluble in water. A solution

of silver nitrate—which is made from silver and nitric acid—is a *test for a chloride*, because it produces a white curdy precipitate with every solution which contains a chloride. A *precipitate* is an insoluble solid which results from a chemical action in solution.

## QUESTIONS FOR STUDY

1. In preparing hydrogen chloride, what substance furnishes the hydrogen? what substance the chlorine?
2. Why is the aqueous solution of hydrogen chloride more useful than the dry gas?
3. How would you prove that a given solution contained a chloride?
4. Why are the methods of preparation used in the laboratory likely to differ from those used commercially?

**Experiment 12. Preparation and Properties of Hydrochloric Acid.**

**Apparatus.** Flask 250 c. c. with stopper carrying thistle tube and delivery tube; ring stand with one ring and one clamp; wire gauze with asbestos center; Bunsen burner; 2 test tubes; wide-mouthed bottle.

**Material.** Sodium chloride; sulphuric acid (2 to 1); blue litmus paper; magnesium; zinc.

a. Pour about 20 c. c. of sulphuric acid (2 to 1) into a flask supported on a wire gauze on a ring stand and add about 10 gms. of sodium chloride. Gently rotate the flask so as to mix the acid with the chloride, close the flask with a stopper carrying a thistle tube and a delivery tube arranged for the collection of gas in a dry test tube by downward displacement. *If necessary*, heat the flask with a *small* flame.

*Describe the action in the generator.*

*Of what elements is hydrogen chloride composed?*

*Which of the original materials furnished the chlorine?*

*Which furnished the hydrogen?*

In this particular case, sulphuric acid is used because it does not vaporize below 338°, so that none of it passes off with the hydrogen chloride.

sodium chloride + sulphuric acid  $\longrightarrow$  sodium sulphate + hydrogen chloride

$\left\{ \begin{array}{l} \text{sodium} \\ \text{chlorine} \end{array} \right.$	$\left\{ \begin{array}{l} \text{hydrogen} \\ \text{sulphur} \\ \text{oxygen} \end{array} \right.$	$\left\{ \begin{array}{l} \text{sodium} \\ \text{sulphur} \\ \text{oxygen} \end{array} \right.$	$\left\{ \begin{array}{l} \text{---} \\ \text{---} \end{array} \right.$
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*Why is hydrogen chloride collected by the method used?*

b. Solubility of hydrogen chloride.



Fill a dish with water and set it on the table. Take the test tube of gas collected by the downward displacement of air, close its mouth tightly with the thumb, invert the test tube and hold its mouth below the surface of the water. Remove the thumb.

*Result?*

*Explain why the gas is not collected over water.*

Close the mouth of the test tube with the thumb and remove it from the water. Moisten a piece of litmus paper with the liquid contained in the test tube.

*Result?*

Taste the liquid.

*Result?*

These effects are typical of the water solution of acids. The solution in the reagent bottle marked "hydrochloric acid" is also prepared by dissolving hydrogen chloride in water.

c. Density of hydrochloric acid.

Pour not more than 10 c.c. of water into a wide-mouthed bottle. Place the mouth of the delivery tube *within* half a centimeter, but *not touching* the surface of the water in the bottle. Heat the flask with a *small flame*, or by a pan of boiling water, for at least ten minutes. While doing this, occasionally look through the water in the bottle horizontally.

*Is the solution of hydrogen chloride formed heavier or lighter than water? Explain.*

d. Action of hydrochloric acid with metals.

Pour half of the solution just made into a test tube and drop into it a strip of magnesium. Bring a flame near the mouth of the tube.

*Result?*

Place a piece of zinc in another test tube and pour the remaining hydrochloric acid upon it. Test the gas with a flame.

*Result?*

*What substance is liberated when hydrochloric acid reacts with these metals?*

Magnesium and zinc are elements.

*When hydrochloric acid reacts with these metals, where does the substance that is liberated come from?*

*What becomes of the metal?*

*What three properties have been mentioned as characteristic of acids?*

Make a drawing of your apparatus, part c.

## CHAPTER X

### GASES AND THEIR MEASUREMENT

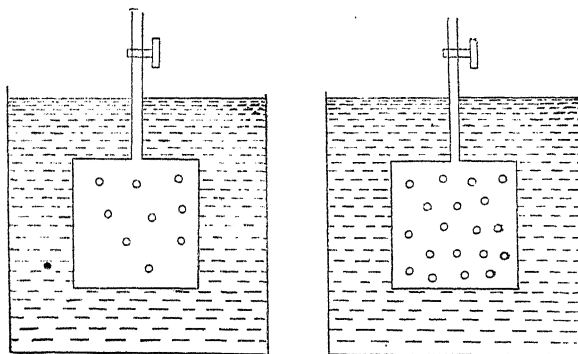
When solids and liquids are heated, in nearly all instances they expand, but the rate of expansion with the temperature is different for each one. The same holds true for the compressibility of liquids and solids. However, when we investigate the behavior of gases with respect to changes in temperature and pressure, we find that the effect on the volume is almost the same no matter what gas is being examined. One does not need to know the name of the gas in order to predict how its volume will change under the influence of changing temperature or pressure.

The fundamental difference between gases on the one hand, and liquids and solids on the other, is the *tendency of gases to expand indefinitely*, so as to fill completely any vessel in which they are contained. If a quantity of gas is introduced into an empty vessel, it distributes itself almost instantly throughout the whole space at its disposal, showing a tendency to still further expansion by exerting a pressure on all parts of the walls of the vessel.

If the vessel already contains one gas, and another gas is admitted, it will be found, after allowing sufficient time to elapse, that both gases are equally distributed throughout the vessel. For example, if ammonia is exposed to the air in one part of a room in which air is perfectly still, a person in a different part of the room will soon be able to smell the ammonia. This process whereby one substance works its way through another, is called *diffusion*.

Now let us ask the question, What sort of structure must we think of a gas as possessing in order to account for these phenomena? The most satisfactory picture is that furnished by the *Kinetic Theory*, which assumes that a gas is composed of particles called *molecules*, whose size is very small compared with the distance between them, and that each molecule is in rapid motion, colliding with other molecules and with the walls of the containing vessel. It is these impacts which produced the observed pressure, a series of rapid impacts

obviously having the effect of a steady pressure. The further assumption is made that the average velocity of the molecules, and hence the pressure produced on the walls of the vessel, depends upon the temperature, increasing as the latter increases. The molecules are in such rapid motion and are most of the time so far apart relatively, that their attractions for each other have very little effect.

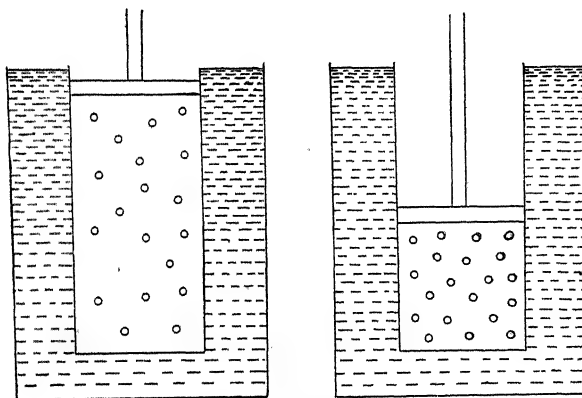


From Hildebrand's "Principles of Chemistry"

Fig. 23. Circles to Represent Molecules

When temperature and volume are kept constant the pressure of a gas is proportional to the amount present.

This picture of the condition of gases has proved to be of such immense service in explaining and predicting their properties that it is hardly questioned by scientists at the present time.



From Hildebrand's "Principles of Chemistry"

Fig. 24. Circles to Represent Molecules

When the temperature and amount of gas are constant the pressure is inversely proportional to the volume.

The atmosphere which surrounds us is a gaseous medium made up mainly of a mixture of oxygen and nitrogen. We are really immersed in an ocean of air, as it were, an ocean which extends for two miles above this earth. We cannot say that a vessel is *empty* unless it were a vacuum, because air fills every little conceivable space. The atmosphere then, being composed of gases, follows the same natural laws which every other gas follows.

The two fundamental gas laws are known as Charles's Law and Boyle's Law.

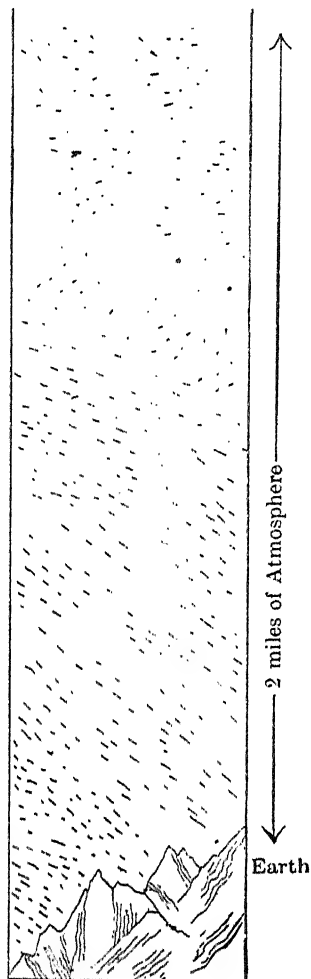


Fig. 25. We are immersed in an ocean of air.

Centigrade and Absolute Scales

Tin	melts
231.9° C.	504.9° A.
Water	boils
100° C.	373° A.
Phosphorus	melts
44° C.	317° A.
Water	freezes
0° C.	273° A.
Mercury	freezes
-38.7° C.	234.3° A.
Chlorine	freezes
-101.6° C.	171.5° A.
Oxygen	boils
-182.9° C.	90.1° A.
Absolute	Zero
-273° C.	0° A.

Fig. 26. Comparison of Temperatures on the Centigrade and Absolute Scales

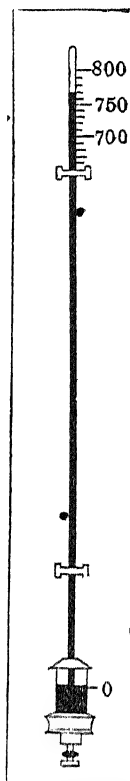


Fig. 27. Mercurial Barometer for Measuring Atmospheric Pressure

Charles's Law states that "*under constant pressure the volume of a gas is proportional to its absolute temperature.*" The centigrade zero (the point at which water freezes into a solid) is equivalent to 273° on

the absolute scale, and the absolute zero then would be  $-273^{\circ}$  centigrade. The volume of the gas would expand  $\frac{1}{273}$  of its volume at  $0^{\circ}$  centigrade for an increase of every degree of temperature, and would contract a corresponding volume for a decrease of each degree of temperature. For example, 273 c. c. at  $0^{\circ}$  C. becomes 274 c. c. at  $1^{\circ}$  C., 283 c. c. at  $10^{\circ}$  C., 303 c. c. at  $30^{\circ}$  C., 253 c. c. at  $-20^{\circ}$  C., etc. All centigrade readings may be changed to absolute readings by adding 273 to them.

Boyle's Law states that "*at a constant temperature the volume of a gas is inversely proportional to the pressure*" the gas supports. According to Boyle's Law, a mass of gas having a volume of 30 c. c. at 700 mm. thus occupies 27.63 c. c. at 760 mm. and 38 c. c. at 600 mm.

$$\text{at } 760 \text{ mm. } v = 30 \times \frac{700}{760} = 27.63 \text{ c. c.}$$

$$\text{at } 600 \text{ mm. } v = 30 \times \frac{760}{600} = 38 \text{ c. c.}$$

If the pressure *increases*, the final volume will be *less*; if the pressure *decreases* the final volume will be *greater*.

Zero degree centigrade and 760 millimeters pressure are known as "standard conditions of temperature and pressure," and volumes of gases are usually compared under these conditions. Seven hundred sixty millimeters represents the pressure or weight of the atmosphere at sea level and is measured by a column of mercury which balances the weight of the air. If the pressure of the atmosphere is greater, more mercury is forced up the tube; if less, the mercury in the barometer would correspondingly fall.

#### QUESTIONS FOR STUDY

1. What will be the volume at  $0^{\circ}$  C. of a mass of gas which at  $20^{\circ}$  C. has a volume of 102 c. c.?
2. One liter (1,000 c. c.) of air at  $15^{\circ}$  C. will have what volume at  $25^{\circ}$  C.?
3. What will be the volume at 800 mm. of a quantity of nitrogen which at 730 mm. measures 1.2 liters?
4. Ninety c. c. of air at  $20^{\circ}$  C. become, when heated, 150 c. c. To what temperature was the air heated?
5. A gas measures 20 liters at  $10^{\circ}$  C. and 760 mm. pressure. What volume will it have at  $30^{\circ}$  C. and 720 mm.?

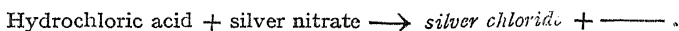
**Experiment 13. Test for a Chloride.**

**Apparatus.** 6 test tubes ; test-tube rack.

**Material.** Dilute hydrochloric acid ; dilute nitric acid ; solutions of silver nitrate ; ammonium hydroxide ; sodium phosphate , potassium oxalate ; 3 unknowns.

a. To a little hydrochloric acid in a test tube add silver nitrate solution. The substance which separates out is silver chloride.

*Complete the equation :*



Any solid which thus separates out of a clear liquid is known as a *precipitate*.

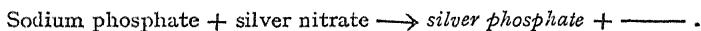
*Describe the precipitate as to color and appearance.*

Set aside the tube containing the precipitate.

b. Take a little of a solution of sodium phosphate and add to it silver nitrate solution.

*Result ?*

*Complete the equation :*



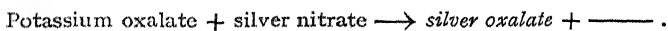
*Describe the color and appearance of the silver phosphate.*

Set the tube aside.

c. Take a little of a solution of potassium oxalate and add to it silver nitrate solution.

*Result ?*

*Complete the equation :*



*Describe the color and appearance of the precipitate.*

Set the tube aside.

d. Add nitric acid to each of the three tubes containing the three precipitates.

*What happens in each case ?*

*How can you distinguish silver chloride from silver phosphate and from silver oxalate ?*

*How can silver chloride be obtained from any soluble chloride ?*

*What two steps would be necessary to distinguish any soluble chloride from any soluble phosphate or oxalate?*

This procedure serves as a means of distinguishing chlorides from other salts.

*Describe a test for a chloride.*

e. Make another portion of silver chloride. Let the precipitate settle, drain off most of the clear liquid, and add ammonium hydroxide.

*Result?*

f. Make still another portion of silver chloride and see what effect light has on it. If possible, stand the tube in the direct sunlight.

What further characteristics of silver chloride are shown in parts e and f?

g. Apply to the instructor for an "unknown." Test it for a chloride, writing the results obtained in each step.

State your opinion as to whether the unknown contained a chloride, and give reason for your belief.

TABLE

NO. OF UNKNOWN	EFFECT OF ADDING SILVER NITRATE	IF PRECIPITATE OBTAINED, EFFECT OF NITRIC ACID ON IT

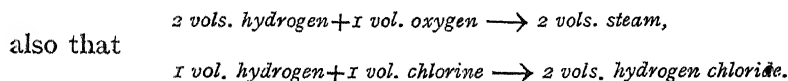


## CHAPTER XI

### MOLECULAR COMPOSITION

We have already learned that all gases behave alike as regards the effects of pressure. Thus, *the volume of any gas varies inversely as the pressure if the temperature remains the same*, according to Boyle's Law. We have also learned that all gases behave alike as regards the effect of temperature. Thus, *the volume of any gas varies directly as the absolute temperature if the pressure remains the same*, according to the Law of Charles.

We learned from the synthesis by volume of steam that



A careful study of the process of making ammonia gas by the direct union of hydrogen and nitrogen gives these results ;



These experimental facts all illustrate the *Law of Gay-Lussac* :

*The volumes of gases used and produced in a chemical change can always be represented by the ratio of small whole numbers.*

These striking cases of uniformity in the behavior of gases led the famous Italian Avogadro, in 1811, to the following explanation. He assumed that *equal volumes of gases under like conditions of temperature and pressure contain the same number of molecules*. In other words, 1 cubic centimeter of hydrogen has just as many molecules as 1 cubic centimeter of oxygen, nitrogen, chlorine, hydrogen chloride, or any other gas, provided the conditions of temperature and pressure are the same.

It should be noticed at once that this theory makes no assumption as to the *actual* number of molecules in a cubic centimeter. But it does give us a simple method of comparing the weights of molecules of gas by comparing the densities of the gases, that is, their weights per unit volume.

We can best show this in a table :

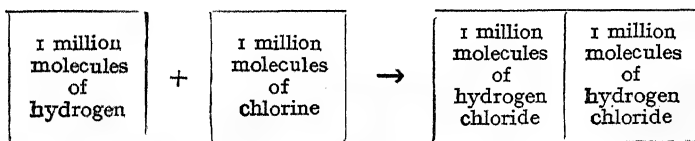
GAS	DENSITY (GRAMS PER LITER)	RELATIVE WEIGHTS OF MOLECULES
Hydrogen	0.09	1
Oxygen	1.43	16
Nitrogen	1.25	14
Hydrogen Chloride	1.64	18.2

In this table we have used only the approximate values for densities and have taken hydrogen as the basis of comparison. It will be seen that the third column gives us the relative weights of the molecules referred to the hydrogen molecule as a standard.

Another very interesting conclusion that we draw from Avogadro's theory is that *every molecule of hydrogen has at least two atoms*. Let us see why this so. Our experiment in the synthesis of hydrogen chloride showed that

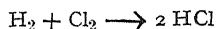


According to Avogadro we must assume that we have the same number of molecules of hydrogen as of chlorine and *twice* the number of molecules of hydrogen chloride. Suppose we have a million molecules of hydrogen, then we must have a million molecules of chlorine and two million molecules of hydrogen chloride. Each one of these molecules of hydrogen chloride must contain at least one atom of hydrogen; so that we have in all *two million atoms* of hydrogen coming from *one million molecules* of hydrogen. This experiment indicates that *each molecule of hydrogen contains at least two atoms of hydrogen*. It may help us to picture this relation if we draw squares to represent the *equal* volumes, as shown below :



By a similar course of reasoning it follows that the chlorine molecule also contains at least two atoms.

We represent these facts by the following equation :



This means that one molecule of hydrogen, containing 2 atoms, plus one molecule of chlorine, containing two atoms, gives two molecules of hydrogen chloride, each containing one atom of hydrogen and one atom of chlorine.

It should be remembered that any *even* number might be used instead of two ; but since there is no reaction in which the hydrogen molecule appears to be divided into more than two parts, we decide that there must be *only two* atoms in the molecule.

From a consideration of the volumetric synthesis of steam, we come by similar reasoning to the conclusion that the oxygen molecule contains two atoms ; and from a study of the synthesis of ammonia, we conclude that the nitrogen molecule also contains two atoms.

#### Experiment 14. Weight of a Liter of Oxygen.

**Apparatus.** Horn pan balance; weights; test tube with rubber stopper carrying delivery tube; sink; 2-liter bottle; glass plate; graduate; ring stand; clamp; Bunsen burner.

**Material.** Well-dried potassium chlorate; dry powdered manganese dioxide; glass wool.

a. Mix about 6 grams of *thoroughly dried* manganese dioxide with 8 grams of potassium chlorate.

Put the mixture in a dry test tube, and above the powder place a loose plug of glass wool.

Tabulate your data as indicated in the table below.

b. Weigh to a centigram the test tube containing the mixture, and the connecting tube.

*Record the weight in the table.*

Clamp the apparatus in place, and adjust the rubber stopper carrying the delivery tube leading to the sink. Collect the gas in a bottle of about 2 liters capacity.

## MOLECULAR COMPOSITION



The test tube should be *inclined at a slight angle* so as to permit the spreading of the black mixture along the tube. Heat with a *small* flame, beginning at the top, and gradually working downward.

*Carefully regulate the heat* so that you can always count the bubbles of the oxygen. Continue the heating until the bottle is nearly full, or as long as time will allow. When the oxygen has ceased to pass over, at once remove the delivery tube. *Do not let cold water suck back into the tube.* Allow the apparatus to cool.

c. While waiting for this, measure the volume of the oxygen. Lower the large bottle into the sink so as to adjust the water levels to the same height, close the mouth of the bottle with a glass plate, and remove the bottle from the sink, inverting it as you do so. *Be careful not to lose any water from your bottle.*

Find the volume of the oxygen by the amount of water necessary to fill the bottle from a graduate.

*Record in the table.*

d. When the test tube has cooled until it feels barely warm to the hand, weigh, as before, the test tube and its contents and the connecting tube.

*Record the weight.*

The loss of weight is the weight of the oxygen evolved.

Record the barometric pressure and the temperature of the water in the sink. This is approximately the temperature of the gas.

TABLE

Weight of generating tube before heating	g.
"    "    "    "    after    "	g.
"    "    oxygen evolved	g.
* Volume of oxygen evolved under conditions of experiment	c. c.
Temperature of oxygen	°C.
Barometric pressure	mm.
Aqueous tension	mm.
Pressure of oxygen	mm.
Volume of oxygen at standard conditions	c. c.
Weight of one liter of oxygen at standard conditions	g.

## CHAPTER XII

### ATOMIC AND MOLECULAR WEIGHTS

We may determine the relative weights of the molecules by comparing the densities of two gases. (The density is the number of units of weight that occupy a unit volume, usually grams per cubic centimeter.) The chemist takes as his standard the *hydrogen atom*, therefore the *molecular weight of a substance is a number which expresses how many times its molecule is heavier than the hydrogen atom*.

For example, the density of oxygen is approximately 1.43 grams per liter, and the density of hydrogen is about 0.09 grams per liter; therefore the density of oxygen is  $\frac{1.43}{0.09}$  or about 16 times the density of hydrogen. According to Avogadro, these densities will have the same ratio as the weights of the molecules. As one molecule of hydrogen contains two atoms, therefore one molecule of oxygen weighs  $2 \times 16$ , or 32, times as much as one atom of hydrogen. In other words, *the molecular weight of oxygen is 32*.

In the same way we can find the molecular weight of nitrogen. Since nitrogen is 14 times as heavy as hydrogen, its molecule is  $2 \times 14$ , or 28, times as heavy as the hydrogen atom. Also, in comparing the densities of hydrogen chloride and hydrogen, we find that hydrogen chloride is 18.25 times as heavy as hydrogen, and therefore its molecular weight is  $2 \times 18.25$ , or 36.5.

We have just seen that the molecular weight of oxygen is 32. The chemist often has occasion to use as many grams of a gas as there are

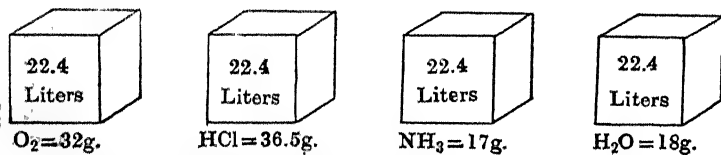


Fig. 28. The Weight of 22.4 Liters of Various Gases

units in its molecular weight, and he calls this weight the *gram-molecular weight*. Thus, 32 grams is the gram-molecular weight of oxygen; 28,

the gram-molecular weight of nitrogen; and 36.5, the gram-molecular weight of hydrogen chloride.

What volume does 32 grams of oxygen occupy under standard conditions? Since one liter of oxygen weighs approximately 1.43 grams, it follows that it takes  $\frac{32}{1.43}$ , or 22.4, liters of oxygen to weigh 32 grams. If we divide the gram-molecular weight of *any gas* by its density, we always get 22.4 liters. Therefore we conclude that the *gram-molecular weight of any gas occupies a volume of 22.4 liters*.

This gaseous volume of 22.4 liters is one of great importance to the chemist in the laboratory, for it furnishes him a quick means of calculating the molecular weight of any substance which can be weighed in the gaseous state; all he has to do is to compute from the weight of a given volume under standard conditions what 22.4 liters would weigh in grams.

Another result of Avogadro's theory is that it gives us a direct method of calculating the relative weights of the atoms, that is, the atomic weights. We have just seen that the oxygen molecule is approximately sixteen times as heavy as the hydrogen molecule, and we have also seen that the oxygen and hydrogen molecules each have two atoms. Therefore we can say that the oxygen atom is sixteen times as heavy as the hydrogen atom; in other words, the *atomic weight of oxygen is 16, referred to the hydrogen atom as a standard*.

In the same way we can compute the atomic weight of nitrogen as 14 and that of chlorine as 35.5.

Unfortunately, we cannot apply this same method to all the elements because we cannot obtain them in the gaseous condition.

We have seen that the molecular weight of a substance expresses the number of times its molecule is heavier than the hydrogen atom; and that the atomic weight of a substance expresses the number of times its atom is heavier than the hydrogen atom, the unit in each case being the same, namely, the hydrogen atom. It is evident, then, that the molecular weight of a compound is equal to the sum of the atomic weights of its constituent atoms.

It is possible experimentally (by means of its density, if a gas) to determine the molecular weight of a compound; it is also possible to determine experimentally its percentage composition by weight. If,

then, we multiply the molecular weight of the compound by the percentage of each element present, we shall get numbers which are the atomic weights, or some multiple of them, of the constituent elements. Further, if we analyze several compounds of the same element, we shall find that the numbers obtained by multiplying the molecular weights by the percentage composition of the element will in each case be the same, or multiples. We then take as the atomic weight the smallest weight of any given element in the compounds considered. In other words, by this method, *the smallest weight of an element found in the gram-molecular weights of its compounds is the atomic weight of that element.*

If we analyze hydrogen chloride we shall find that 35.5 grams of chlorine are united with 1 gram of hydrogen. The chemist calls this number 35.5 the *combining weight* or *reacting weight* of chlorine. The *combining* or *reacting weight of an element is the weight of that element which combines with or replaces 1 gram of hydrogen.* It will be seen at once that this combining weight is exactly equal to the atomic weight of chlorine, which we have already determined by comparing the densities of the elementary gases chlorine and hydrogen. Indeed, this combining weight *must* be its atomic weight, because a molecule of hydrogen chloride consists of but one atom of hydrogen combined with one atom of chlorine.

Unfortunately, the combining weights of some elements are *not* equal to their atomic weights. If we analyze water we find that one part of hydrogen unites with eight parts by weight of oxygen, and therefore the combining weight of oxygen is 8, or *one half its atomic weight.* It has been found in general that *the atomic weight of an element is equal to its combining weight or some multiple of it.* In the case of oxygen this multiple is 2.

Let us take the case of an element which does not combine with hydrogen but which replaces it. Such an element is sodium; its combining weight may be determined as follows:

If we put metallic sodium in water, hydrogen and sodium hydroxide are formed. When this experiment is done quantitatively, we find that 23 grams of sodium replace 1 gram of hydrogen. Therefore the combining weight of sodium is 23.

## PROBLEMS

1. When 150 c. c. of hydrogen burn in chlorine, how many cubic centimeters of hydrogen chloride are produced?
2. If the molecular weight of nitrous oxide is 44, what is the approximate density (grams per liter) of the gas?
3. If 503 c. c. of marsh gas under standard conditions weigh 0.36 grams, what is the molecular weight of marsh gas?
4. The analysis of ammonia gas shows that it is 82.4% nitrogen, and the remainder hydrogen. Compute the combining weight of nitrogen.

**Experiment 15. Equivalent of Magnesium.**

**Apparatus.** Gas-measuring tube; 50 c. c. dish for water; thermometer; barometer.

**Material.** Magnesium; concentrated hydrochloric acid; thread.

Roll a piece of magnesium whose exact weight is known into a loose coil somewhat smaller than the inside diameter of the gas-measuring tube. Pass a thread through the loop of the coil and tie it.

Pour about 5 c. c. of concentrated hydrochloric acid into the gas-measuring tube. Fill the remainder of the tube with water, taking care not to mix the acid and water; the heavier acid will remain at the bottom.

Put the magnesium into the measuring tube, holding the thread so that the magnesium will not sink. Close the tube with the thumb, and invert it into a dish of water, resting the mouth of the tube against the bottom so that the thread will be held, allow the magnesium to rise not quite to the first marking in the tube.

The heavier acid will flow down and react with the magnesium.

*What is the gas that collects?*

*From what substance does it come?*

*Complete the equation:*

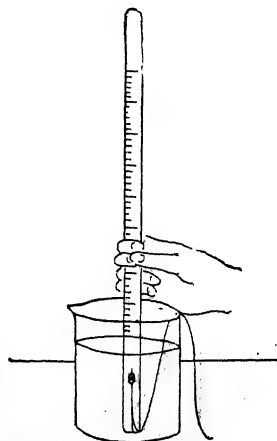
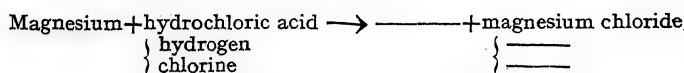


Fig. 29. Collecting Gas to Be Measured in a Gas-Measuring Tube



*Why is this reaction called a replacement?*

When the action has ceased (all the metal being dissolved), measure the difference in levels.

*Read the volume of hydrogen obtained and record your result.*

*Record the temperature of the liquid used and the barometric pressure. Correct the pressure for aqueous tension. Make all calculations in the table.*

*Reduce the volume of hydrogen to standard conditions.*

The weight of 1,000 c. c. of hydrogen is 0.09 g.

*Calculate the weight at standard conditions of the hydrogen that was produced by the reaction using the proportion:*

$$1,000 \text{ c. c. of hydrogen : corrected volume :: } 0.09 \text{ gm. : } x \text{ gm.}$$

This result is the weight of hydrogen that is liberated by the action of your known weight of magnesium.

*Calculate how much magnesium would have been necessary to liberate 1 gram of hydrogen.*

This weight is called the equivalent of magnesium.

*The equivalent of any element is the weight of that element that replaces (or combines with) 1 gram of hydrogen.*

TABLE

Weight of magnesium taken	g.
Volume of hydrogen obtained	c. c.
Temperature	° C.
Barometric pressure	mm.
Aqueous tension	mm.
Corrected pressure	mm.
Volume of hydrogen under standard conditions	c. c.
Weight of hydrogen (calculated)	g.
Equivalent of magnesium (calculated)	g.

## CHAPTER XIII

### SYMBOLS AND FORMULAS

A symbol is simply the initial letter, or the initial letter followed by another characteristic letter of the name of the element. Each symbol stands for *an atom* of the element. Thus, the symbol O stands for one atom of oxygen and the symbols H, N, Na, Fe, represent *one atom* of hydrogen, nitrogen, sodium, and iron, respectively, and not these substances in general. More exactly then, they are *atomic symbols*, and not merely abbreviations of the elements.

We have also seen that the *atoms have a perfectly definite weight*, and so the symbol means not only the element but a definite quantity of the element. O always means 16 parts by weight of oxygen, and C always means 12 parts by weight of carbon.

When we know how many atoms of an element go to make up a molecule of that element, we can express this knowledge very briefly by symbols with small subscripts placed after them. We have already seen that the molecule of hydrogen, oxygen, nitrogen, chlorine, and some others consist of 2 atoms; this fact is represented by the symbols  $H_2$ ,  $O_2$ ,  $N_2$ ,  $Cl_2$ , respectively. The symbol  $H_2$  thus means *one molecule of hydrogen*.

We may represent the composition of compounds with these symbols. This is done simply by placing side by side the symbols for the various elements in the compound. *Symbolic expressions for compounds enable us to see at once what elements make up the compounds.* Thus, while the common names of salt, water, galena, and caustic soda give us no idea of the composition of these bodies, the corresponding formulas  $NaCl$ ,  $H_2O$ ,  $PbS$ , and  $NaOH$  do. (The arrangement of symbols representing a compound is called the formula of that compound.) Common salt, sodium chloride, is made up of one atom of sodium (Na) and one atom of chlorine (Cl) and is represented by the united symbols of the two elements  $NaCl$ .

When the molecules of a compound contain more than one atom of any particular element, this fact is expressed by placing small numbers

immediately after the symbol of that element. Thus,  $\text{H}_2\text{O}$  is the formula for one molecule of water, a compound containing two atoms of hydrogen and one atom of oxygen. Very often it is necessary to indicate a group of atoms in a molecule which behave in many reactions like a single atom. Parentheses are used for this purpose. Thus, a molecule of calcium hydroxide (lime water) may be represented by  $\text{Ca}(\text{OH})_2$ , which indicates a molecule containing one atom of calcium, two atoms of oxygen, and two atoms of hydrogen.

When we wish to indicate more than one molecule of any substance, we place large numerals before the symbol or formula. Thus  $2\text{O}_2$  means 2 molecules of oxygen and  $5\text{H}_2\text{O}$  means five molecules of water.

If we know the formula of a compound and have a table of atomic weights, we can easily calculate the percentage of each element in the compound. Given the formula of potassium chlorate as  $\text{KClO}_3$ , and the atomic weight of potassium, 39, of chlorine, 35.5 and of oxygen, 16, then the formula represents

39.0 parts by weight of potassium,  
35.5 parts by weight of chlorine,  
and  $3 \times 16$ , or 48, parts by weight of oxygen.

Therefore potassium chlorate contains  $\frac{39}{122.5}$ , or 0.318, or 31.8 per cent of potassium,  $\frac{35.5}{122.5}$ , or 0.290, or 29 per cent of chlorine, and  $\frac{48}{122.5}$ , or 0.392, or 39.2 per cent of oxygen.

### QUESTIONS FOR STUDY

1. Give the symbol of each of the following elements: sodium, sulphur, manganese, magnesium, mercury, potassium, phosphorus, lead, platinum, and iron.
2. Name the elements which correspond to the following symbols: Ag, Sn, Ni, Au, Cu, Ra, Al, Ba, P, Cl, Hg.
3. State the facts expressed by the formulas,  $\text{CuSO}_4$ ,  $3\text{NaCl}$ ,  $2\text{H}_2\text{O}$ ,  $4\text{P}$ , and  $\text{Ba}(\text{OH})_2$ .
4. Compute the percentage composition of the following compounds from the formulas as given: (a) common salt ( $\text{NaCl}$ ); (b) saltpeter ( $\text{KNO}_3$ ); (c) mercuric oxide ( $\text{HgO}$ ); (d) sulphuric acid ( $\text{H}_2\text{SO}_4$ ); and (e) cane sugar ( $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ ).

### Experiment 16. Heating of Metals in Air.

**Apparatus.** Bunsen burner; forceps; ring stand with one ring; pipe-stem triangle; lid of porcelain crucible; iron wire, 15 cm. long.

**Material.** Copper wire; magnesium ribbon, 6 cm.; granulated tin; sandpaper.

a. Copper.

Scour a piece of copper with sandpaper. Examine the bright copper, noting its color, luster, and flexibility. Take hold of one end of the copper with forceps, and hold the other end in the outer flame of the burner till it is red-hot. Remove the strip from the flame and watch it while cooling. Bend the strip.

*Compare the properties of the surface material with those observed in the original copper and record in the table below.*

b. Magnesium.

Examine a piece of magnesium ribbon, noting its color, luster, and flexibility. Using forceps, take hold of one end of the magnesium and place the free end of the ribbon in the flame.

*Result?*

*Compare the product with the magnesium and record your observations in the table.*

c. Tin.

Place the lid of a porcelain crucible on a pipe-stem triangle, supported on a ring stand. On the crucible lid put a few pieces of granulated tin and heat gently at first, keeping the flame in motion and *well below the crucible lid*. When the tin melts, stand the burner beneath the crucible lid and stir the tin constantly with an iron wire.

*Compare the product with the original tin and record your observations in the table.*

*Have chemical or physical changes taken place during the heating of the metals in air? Explain.*

TABLE

MATERIAL EXAMINED	COLOR	LUSTER	FLEXIBILITY
Copper			
Substance obtained by heating copper			
Magnesium			
Substance obtained by heating magnesium			
Tin			
Substance obtained by heating tin			

**Experiment 17. Types of Chemical Change: Direct Combination—Simple Decomposition.**

**Apparatus.** Hard glass test tube; Bunsen burner; two 4-in. test tubes; forceps; ring stand.

**Material.** Iodine; yellow phosphorus; filter paper; mercuric oxide; wooden splinter.

*Relation of energy to chemical action.*

Energy is the ability to do work. The more common ways in which it shows itself are through heat, light, and electricity.

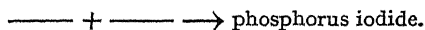
Name chemical reactions in which these forms of energy are concerned.

*a. Direct combination, or synthesis.*

Put a small, thin piece of phosphorus on a piece of filter paper on your iron ring stand. *Phosphorus must be kept under water until the moment of use. It takes fire easily and its burns are serious.* With forceps, lay a crystal of iodine on the piece of phosphorus. *Stand well back, as the action is vigorous.*

*Are phosphorus and iodine elements or compounds?*

Complete the equation:



*Why is the reaction called a direct combination?*

*Is energy liberated or absorbed in this reaction?*

*State reasons for your answer.*

*Name three other experiments you have performed in which elements combined directly without continuous application of heat.*

State in each case whether, after the action was once started, energy was absorbed or liberated.

This evidence will guide you in choosing a word to insert in the blank in the statement of the following important principle:

*Direct combination takes place readily only when energy is ———.*

*b. Direct decomposition.*

Heat a little mercuric oxide in a hard glass test tube. Test for oxygen from time to time

*What forms on the sides of the tube?*

*Write an equation for the reaction.*

*Why is the reaction called a direct decomposition?*

Stop heating the mercuric oxide.

*Does the action continue?*

*Is energy absorbed or liberated in this action?*

The instructor will have in operation, during the laboratory period, electrolysis apparatus in which an electric current is passing through a concentrated solution of hydrochloric acid. The gas that is given off from the anode is collected over a saturated solution of common salt.

*What gas is given off at the anode? at the cathode?*

Stop the flow of the current.

*Does the action continue?*

*Is energy liberated or absorbed in this reaction?*

*Write an equation for the reaction.*

*Why is it called a direct decomposition?*

The evidence of these two experiments will guide you in choosing a word to fill in the blank in the statement of the following principle:

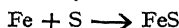
*In most cases of direct decomposition, energy is ——— .*

## CHAPTER XIV

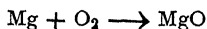
### CHEMICAL EQUATIONS

The principal use which the chemist makes of symbols and formulas is in writing in a condensed and precise form a large amount of information about chemical changes or reactions. In as much as matter is indestructible (Law of Conservation of Matter) *we represent these changes or reactions as equations*. There is the same amount after the change as before.

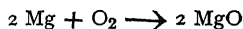
The symbols and formulas of *all the materials undergoing change* are placed on the *left*; those of the new *products* resulting from the change on the *right*; the arrow means *forms* or *gives*. For example, the union of iron and sulphur to form iron sulphide is expressed thus:



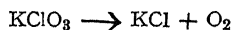
The burning of magnesium in oxygen might at first be expressed thus:



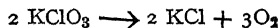
but this equation is *wrong* because we have *two* atoms of oxygen on the left and *only one* on the right side, and this would indicate that some of the oxygen had been lost. The very word "equation" means that there must be just as many atoms of each element after the change (on the right) as before (on the left). We can "balance" the equation, as it is called, thus:



In decomposing potassium chlorate by heat we might write the equation at first thus:



and then balance it thus:



From a study of the above it will be seen that these equations are not algebraic; they represent chemical changes which actually take place. Before we can write an equation we must know:

- (1) *that the change takes place;*
- (2) *what substances are involved;*
- (3) *what the products are; and*
- (4) *the formulas of all the substances involved and of all products.*

In making a chemical equation there are *five* steps:

(1) Write *on the left* the formula of each substance taking part in the reaction.

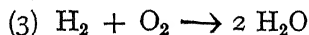
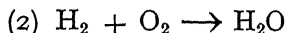
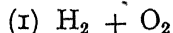
(2) Write *on the right* the formula of each substance known to be produced.

(3) Correct the right-hand side by increasing the number of molecules so as to use all the elements on the left.

(4) Correct the left-hand side by increasing the number of molecules needed by the substances produced.

(5) Check up to make sure that each atom of all elements is accounted for.

*For example*, hydrogen burns in oxygen and produces water.



(5) We have 4 atoms of hydrogen in the 2 molecules of hydrogen on the *left*, and we have 4 atoms of hydrogen in the 2 molecules of water on the *right*. We have 2 atoms of oxygen in 1 molecule of oxygen on the *left*, and we have 2 atoms of oxygen in the 2 molecules of water on the *right*.

*Cautions.* Use the *molecular formulas* for the gaseous elements. Thus the molecular formula of oxygen, hydrogen, nitrogen, and chlorine are  $\text{O}_2$ ,  $\text{H}_2$ ,  $\text{N}_2$ ,  $\text{Cl}_2$ .

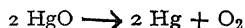
Remember that the number placed before a formula multiplies the *whole* formula. Thus  $2 \text{H}_2\text{O}$  is the same as  $2 (\text{H}_2\text{O})$  which means two molecules of water consisting of 4 atoms of hydrogen and 2 atoms of oxygen.

We balance equations *not* by altering the formulas of the substances, but by taking the proper number of molecules so that the number of atoms of each element is the same on both sides.

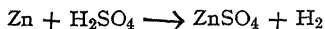


*What an equation does not show.* Chemical changes will usually occur only under certain *conditions* which are not indicated in the equation.

For example, the equation

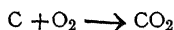


does not show that it is necessary to keep heating the red oxide of mercury in order to decompose it. The equation



does not indicate that the sulphuric acid must be diluted with water.

In general, the materials which are present but which undergo no change, such as water used as a solvent and the various catalysts, are not represented in the equation. Finally, we do not as a rule try to show the amount of heat which is evolved in the chemical reaction. For example,



shows that carbon unites with oxygen to form carbon dioxide; but it does not tell *how much heat is produced*, a piece of information which is most important for the industrial chemist.

### QUESTIONS FOR STUDY

1. Why is a chemical equation called an *equation*?
2. What is the difference between  $\text{O}_2$  and  $2 \text{O}$ ?
3. When zinc and sulphur are heated together zinc sulphide ( $\text{ZnS}$ ) is formed. Write the equation.
4. When aluminium reacts with hydrochloric acid ( $\text{HCl}$ ), hydrogen and aluminium chloride ( $\text{AlCl}_3$ ) are formed. Write the equation.
5. When carbon dioxide is bubbled through limewater [ $\text{Ca}(\text{OH})_2$ ], calcium carbonate ( $\text{CaCO}_3$ ) and water are produced. Write the equation.

### Experiment 18. Types of Chemical Change—Simple Replacement.

**Apparatus.** 2 test tubes.

**Material.** Zinc; dilute hydrochloric acid (1 to 4); iron filings; saturated solution of copper sulphate.

a. Place a piece of zinc in a test tube and add 5 c. c. of dilute hydrochloric acid. After the action has continued for several minutes, feel the test tube.

*Is energy liberated or absorbed in the reaction*

*Write an equation for the reaction.*

*Why is this said to be a case of simple replacement?*

b. To 8 c. c. of a saturated solution of copper sulphate contained in a test tube, add 2 c. c. of fine iron filings. Close the mouth of the tube with the thumb and shake the solution back and forth several times. After the reaction has continued a few minutes, feel the test tube.

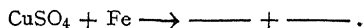
*Is energy liberated or absorbed in the reaction?*

*Allow the test tube to stand for several minutes.*

*What change takes place in the color of the solution?*

*What change takes place in the appearance of the solid?*

Complete the equation:



*Why is this called a case of simple replacement?*

From the evidence of the two experiments, choose a word to fill in the blank in the following statement of an important principle:

*When simple replacement takes place, energy is ---.*

*Heats of formation.*

In the case of most compounds, energy is liberated when the elements unite to form the compound. The amount of energy is known as *the heat of formation*. In the case of certain compounds energy must be furnished as it is absorbed when the compound is formed. In such a case, the substance is said to have a negative heat of formation. The amount of energy is expressed in calories of heat. *A calorie is the amount of heat necessary to warm one gram of water one degree centigrade.* The heat of formation is the number of calories of heat absorbed or liberated during the formation of one gram molecule of a compound from its elements.

The following table gives the heats of formation of certain compounds.

TABLE

*Heats of Formation of Certain Compounds*

	CALORIES
Calcium aluminium silicate	1,195,500
Calcium carbide	-6,250
Carbon disulphide	-19,000
Copper (cupric) chloride (dil. sol.)	62,500
Copper (cupric) sulphate (dil. sol.)	197,500
Hydrogen chloride	22,000
Hydrogen bromide	8,400
Hydrogen iodide	-7,000
Hydrogen sulphide	4,800
Iron carbide	8,460
Iron (ferrous) chloride (dil. sol.)	100,100
Iron (ferric) chloride (dil. sol.)	255,700
Iron (ferrous) sulphate (dil. sol.)	234,900
Iron (ferric) sulphate (dil. sol.)	650,500
Magnesium chloride (dil. sol.)	187,100
Magnesium sulphate (dil. sol.)	321,100
Mercuric chloride (dil. sol.)	50,300
Mercuric cyanide	-62,500
Nitrous oxide	-20,600
Nitric oxide	-21,600
Phosphorus pentoxide	369,400
Potassium iodide (dil. sol.)	81,800
Potassium bromide (dil. sol.)	90,400
Potassium chlorate	93,800
Potassium chloride (dil. sol.)	101,200
Silicon carbide	1,963
Silver oxide	7,000
Sodium chloride (dil. sol.)	96,900
Sodium iodide (dil. sol.)	70,400
Zinc chloride (dil. sol.)	113,300
Zinc cyanide	-27,900

*Bearing in mind the first principle stated in last week's experiment, pick out from the table five compounds that could easily be formed by direct combination.*

*Name three that could not be formed easily by direct combination.*

In decomposing a compound, exactly as much energy must be furnished as is liberated when the compound is formed.

*Bearing this fact in mind, and also the second principle stated in last week's experiment, name five compounds that it would be difficult to decompose.*

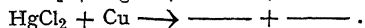
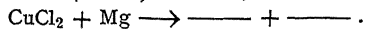
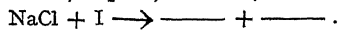
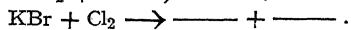
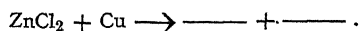
*Also name five that would be easily decomposed.*

*Name three that would decompose with liberation of energy.*

To decide whether or not replacement actions will take place, compare the heats of formation of the original compound with that of the one that might be formed in the reaction.

*What principle, then, will guide you in making your decision?*

*Of the following equations, complete those which you think will actually occur; in the other cases write the words "No reaction."*



Prove your conclusions in one or two cases.

## CHAPTER XV

### CHEMICAL CALCULATIONS

To the industrial chemist the chemical equation is very important because it enables him to *calculate* just how much material he ought to use in a given reaction and how large a product he may hope to get. In actual practice, however, he very seldom gets the full amount of the product as calculated from the equation, and therefore he computes the efficiency of his process. These computations are all based on the chemical equations of the reactions, and require only the simplest kind of arithmetical work.

An equation when properly balanced tells us what substances react and what the products are; it also gives us the number of molecules of each substance involved; finally, *by using the atomic weights, we learn the relative weights of the different substances in the equation.*

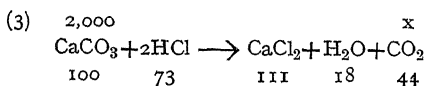
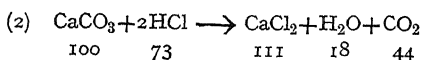
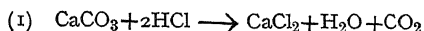
When only weights, or only volumes are involved, the solution of the problem is very simple. *Problems involving weight and volume* may be solved first by finding the weight of the gas involved and thereby computing the volume from the density. Or, if the volume is given, compute the weight and solve by proportion.

Let us sum up in five general rules *the method used in solving chemical problems*:

- (1) Write the complete equation for the reaction in question.
- (2) Write under each formula the figure representing the weight for which each complete formula stands.
- (3) Write above the equation what we know about each of the substances involved in the reaction.
- (4) Read the equation and the figures as they now stand and make a proportion which will show the required value.
- (5) Solve this proportion and decide whether the answer is reasonable by making a rough calculation as to what the answer should be.

Suppose we wish to find out how many liters of carbon dioxide (measured under standard conditions) can be obtained from 2 kilograms of marble by the action of hydrochloric acid.

According to the five general rules, we have



$$(4) 100 : 44 :: 2,000 : x$$

$$(5) x = \frac{44 \times 2,000}{100} = 880 \text{ grams.}$$

Finally, we look up the density of carbon dioxide and find it to be about 1.98 grams per liter.

Therefore the volume  $\frac{880}{1.98} = 444$  liters.

### PROBLEMS

1. What weight of zinc will be needed to generate 100 liters of hydrogen, using dilute sulphuric acid?

2. Calculate the weight of hydrogen generated when 100 c. c. of dilute sulphuric acid containing 25 grams of pure sulphuric acid is added to 10 grams of zinc.

3. What volume of hydrogen (measured under standard conditions) could be obtained by the electrolytic decomposition of 5 grams of water?

4. What volume of carbon dioxide (measured under standard conditions) would be given off from 1 kilogram of sodium bicarbonate in a fire extinguisher? The equation is  $2\text{NaHCO}_3 + \text{H}_2\text{SO}_4 \longrightarrow 2\text{CO}_2 \uparrow + 2\text{H}_2\text{O} + \text{Na}_2\text{SO}_4$ .

### Experiment 19. Types of Chemical Action: Double Decomposition—Insolubility.

**Apparatus.** 6 test tubes.

**Materials.** Solutions of barium nitrate, lead nitrate, silver nitrate, ammonium chloride, copper sulphate (all approximately N/5), sodium hydroxide (1 to 10), dilute hydrochloric acid (1 to 3).

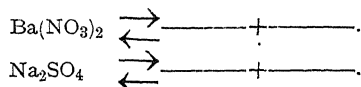
*Effect of insolubility of one of the products.*

To 5 c. c. of a solution of barium nitrate, add a little of a solution of sodium sulphate.

*What evidence is there that chemical action has occurred?*

Allow the tube to stand a few minutes. The substance that settles is barium sulphate.

Show by completing the following equations what ions are formed by barium nitrate and by sodium sulphate.



Explain why these actions are reversible actions.

Show by an equation how barium sulphate results on mixing the two solutions.

In this case, is there equilibrium, or an action that goes to an end?

Explain why such an action is called a double decomposition.

The remaining parts of the experiment are intended to illustrate the conditions under which, on mixing two substances that ionize, we get either an equilibrium or an action that goes to an end.

Try the action between solutions of the substances paired in the following table; record your observations of the results in Column 2; the information needed for Column 5 is in Table III of the Appendix; fill in the last column after studying the rest of the table as a whole.

Complete the following sentence:

*A double decomposition will go to an end if one of the possible products is — .*

TABLE

SUBSTANCES WHOSE SOLUTIONS ARE MIXED	DOES ACTION GO TO AN END, OR REMAIN IN EQUILIBRIUM?	IONS EXISTING IN ORIGINAL SOLUTION	PRODUCTS POSSIBLE BY NEW COMBINATION OF IONS	POSSIBLE PRODUCTS SOLUBLE OR INSOLUBLE? SEE TABLE III	REASON FOR ACTION (IN CASE IT GOES TO AN END)
$\text{AgNO}_3 + \text{NaCl}$			(a)	(a)	
$\text{Pb(NO}_3)_2 + \text{Na}_2\text{SO}_4$			(b)	(b)	
$\text{CuSO}_4 + \text{NaOH}$			(a)	(a)	
$\text{NH}_4\text{Cl} + \text{Na}_2\text{SO}_4$			(b)	(b)	
$\text{Pb(NO}_3)_2 + \text{NaCl}$			(a)	(a)	
$\text{Ba(NO}_3)_2 + \text{NaCl}$			(b)	(b)	
$\text{CuSO}_4 + \text{NH}_4\text{Cl}$			(a)	(a)	
			(b)	(b)	

**Experiment 20. Types of Chemical Action: Double Decomposition — Volatility.****Apparatus.** 5 test tubes.**Materials.** Solutions of sodium hydroxide (1 to 10), dilute hydrochloric acid (1 to 3), solid sodium sulphate, copper sulphate, sodium carbonate, ferrous sulphide, ammonium chloride.*Effect of volatility of one of the products.*

In these experiments, use the first-named substance as a solid. Fill the curved bottom of the tube with the substance and add about 2 or 3 c. c. of the second compound. Fill in the table as in Experiment 19. Information for Column 5 may be found in Table V of the Appendix.

TABLE

SUBSTANCES USED	DOES THE ACTION GO TO AN END?	IONS THAT COULD BE FORMED	PRODUCTS POSSIBLE BY NEW COMBINATION OF IONS	POSSIBLE PRODUCTS VOLATILE? SEE TABLE V	REASON IN CASE ACTION GOES TO AN END
$\text{Na}_2\text{SO}_3 + \text{HCl}$			(a)	(a)	
$\text{CuSO}_4 + \text{HCl}$			(b)	(b)	
$\text{Na}_2\text{CO}_3 + \text{HCl}$			(a)	(a)	
$\text{FeS} + \text{HCl}$			(b)	(b)	
$\text{NH}_4\text{Cl} + \text{NaOH}$			(a)	(a)	
			(b)	(b)	

*Complete the following sentence:**Double decompositions go to an end if one of the possible products is ———.*

It will be seen from experiments 19 and 20 that a double decomposition goes to an end if one of the products *leaves the field of action*. If water is formed in a double decomposition, it also leaves the field of action, because water does not form ions readily, and therefore is as much out of the action as if it were insoluble or volatile.

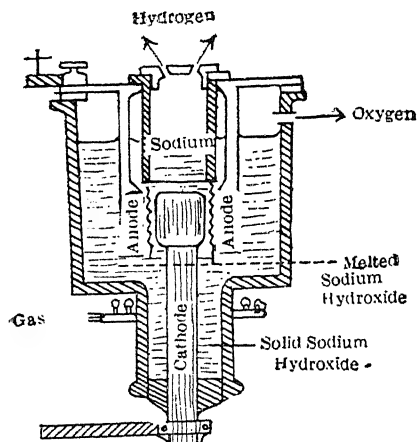
*Explain why a reaction between a base and an acid goes to an end.*



## CHAPTER XVI

### SODIUM AND POTASSIUM

The element sodium is very widely distributed in nature and occurs in combination with other elements in many places in large quantities. On account of its great chemical activity it does not occur in nature in the free condition. Nearly all the salts of sodium are soluble in water. We should, therefore, expect to find most of the sodium compounds dissolved in the sea, and such is the fact. When the rocks



From "Elementary Principles of Chemistry,"  
Brownlee, Fuller, etc.

Fig. 30. It is difficult to separate sodium from its compounds. Electrolysis of sodium hydroxide gives us the best means of accomplishing this result.

undergo weathering and set the sodium compounds free, these dissolve readily in water and are swept down to the sea. In this way, compounds of sodium and especially sodium chloride have been accumulating for ages in the sea, and this is part of the explanation of the saltiness of sea water. Since practically all of the simple salts of sodium readily dissolve in water, we do not find an accumulation of these salts in regions where there is appreciable rainfall. In Chile large beds of sodium nitrate are found which from their analogy to potassium nitrate, or ordinary saltpeter, are known as "Chile saltpeter."

Sodium occurs in small quantities practically everywhere and in everything. We have a very sensitive means in the spectroscope of detecting the presence of minute quantities of sodium. Indeed, the atmospheric air always contains sodium. The universal presence of sodium seems to be due to its existence in the atmosphere. The chloride is taken up with the water vapor over the sea, and distributed in minute quantities everywhere.

On a large scale sodium is prepared by electrolysis of the fused hydroxide.

The element sodium is a soft solid which, when freshly cut with a knife, has a metallic luster and a steel-gray color. The surface becomes quickly tarnished due to the rapidity with which it takes up oxygen from the air or from moisture, forming the oxide or hydroxide. Sodium is a very active substance chemically. It combines readily with moist oxygen, but very slowly with dry hydrogen. With potassium, sodium forms *alloys* which are liquid at ordinary temperatures. Sodium, in the presence of water, forms the hydroxide NaOH. This, we shall learn, is one of the very strongest bases, and as we would expect, combines with all acids. In the light of these facts it is most remarkable that *perfectly dry sodium does not react with perfectly dry sulphuric acid*.

Potassium, like sodium, does not occur in nature in the free condition, and for the same reason, i. e., the great chemical activity of the substance. Although the salts of potassium, like those of sodium, are soluble in water, they are not carried down to the sea in anything like the same relative quantity.

This is due to the fact that plants have the power of taking up potassium compounds in large quantities and building them up in their tissues. They exercise selective absorption for potassium compounds allowing the sodium ones to be carried on by the same waters from which they remove the potassium. When such plants are burned, the potassium salts remain behind in the ashes. Potassium occurs in large quantity in the ashes of certain kinds of wood and can easily be leached out of the ashes by means of water which is allowed to trickle through them. The *lye* thus obtained contains a large amount of potassium hydroxide.

Potassium also occurs in many of the more common rocks and minerals in the form of silicates, and in the great salt beds, especially those of Stassfurt, Germany. The chloride is known as *sylvite*, the nitrate as *saltpeter*, and the sulphate, when in combination with other metallic sulphates, as *alums*.

Metallic potassium is prepared by electrolyzing the chloride or hydroxide.

Potassium is characterized by its great chemical activity, being even more active than sodium. It combines with the oxygen of the air with the greatest readiness, and like sodium must be preserved under petroleum. *Potassium does not combine with dry oxygen* but combines with the greatest vigor if the merest trace of moisture is admitted. This is another example of the wonderful influence exerted by water on chemical activity.

#### QUESTIONS FOR STUDY

1. How are the elements sodium and potassium found in nature? Why?
2. Explain the action of sodium and potassium with water; with dry oxygen; with moist oxygen.
3. Account for the fact that solutions of sodium carbonate and potassium carbonate are basic.
4. What nonmetallic element is obtained from the deposits of Chile saltpeter?

#### Experiment 21. Bases.

**Apparatus.** Evaporating dish; glass plate; stirring rod; four test tubes; glass funnel; ring stand with one ring; Bunsen burner; wire gauze.

**Material.** Metallic sodium; filter paper; red litmus paper; calcium oxide; sodium hydroxide solution about 32 g. per 100 c. c. of water; ammonium hydroxide solution, 1 to 5; hydrochloric acid, 1 to 1; ferric chloride solution, 8 g. per 100 c. c.

*a. Very active metals form bases by direct action with water.*

Take a freshly cut piece of sodium the size of a pea and completely remove the adhering oil with filter paper. Hold a square of glass vertically (to protect the face) in front of an evaporating dish containing about 10 c. c. of water. Place the sodium on the water.

*Describe the action.*

*What gas is liberated?*

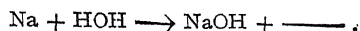
Determine the action of the solution on litmus and rub some of it between the fingers.

*Results?*

Evaporate part of the solution to dryness. The substance left in the dish is sodium hydroxide.

*Describe its appearance.*

*Complete the equation:*



*Is sodium hydroxide a soluble base?*

*b. Some bases can be prepared by action of the oxide of the metal with water.*

Place a gram of quicklime (calcium oxide,  $\text{CaO}$ ) in a test tube and add 1 c. c. of water. Warm the mixture until the action starts; remove the tube from the flame and see if there is continued action between the quicklime and the water. If not, warm the tube until an action begins and continues without the addition of heat.

*Describe the action.*

*How does the substance formed compare with the original lime?*

This new substance is slaked lime (calcium hydroxide,  $\text{Ca(OH)}_2$ ).

*Write an equation for its formation.*

Add water to the slaked lime, shake thoroughly, and allow the solid to settle. Pour off the clear liquid into another test tube, retaining the solid for part *d*. Determine the action of this solution on litmus.

*Result?*

*Is calcium hydroxide a soluble base? State reasons for your answer.*

*What common property have the solutions of sodium hydroxide and calcium hydroxide?*

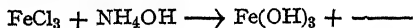
This is one of the characteristic properties of the solutions of metallic hydroxides (bases).

*c. Other bases can be prepared by a process of precipitation.*

To 5 c. c. of a hot solution of ferric chloride add 4 c. c. of ammonium hydroxide.

*Result?*

*Complete the equation:*



Filter the solution and thus obtain the ferric hydroxide on a filter paper. Wash it thoroughly, using three separate portions of water. Allow each portion of the water to drain through *completely* before the next is added.

Transfer the washed precipitate to a test tube and shake the mixture thoroughly. Allow the precipitate to settle and pour off most of the clear liquid. This is to complete the washing of the ferric hydroxide. Add 5 c. c. of water, shake the mixture, and determine its action on litmus.

*Is there any evidence that ferric hydroxide is soluble?*

Retain the contents of the tube for part *d*.

*What characteristic must bases have in order to act on litmus?*

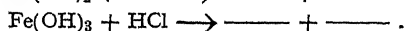
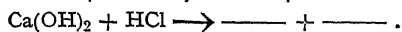
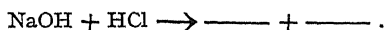
*d. Determination of the action of bases with acids.*

Place in one test tube sodium hydroxide solution, in a second, some of the solid calcium hydroxide from part *b*, and in a third, the ferric hydroxide from part *c*. To each add 5 c. c. of hydrochloric acid. Determine whether heat is produced in each case.

*Results?*

*Do you observe any other evidences of chemical action?*

*Complete the equations.*



TABLE

NAME OF BASE	SOLUBLE OR INSOLUBLE	ACTION ON LITMUS

## CHAPTER XVII

### SOLUTIONS

Solutions of acids, bases, and salts differ conspicuously in certain properties from solutions of other substances (e. g., sugar, glycerine, alcohol). That is, we have two classes of solutions. One class—acids, bases, and salts,—is active chemically, conducts an electric current, and behaves abnormally when boiled or frozen, whereas the other class—not acidic, basic, or salty—is not active chemically, does not conduct an electric current, and behaves normally when boiled or frozen.

The properties of solutions, especially solutions of acids, bases, and salts, are explained more or less acceptably by the *theory of ionization*, which was proposed in 1887 by the Swedish chemist Arrhenius. The theory is usually stated as follows:

*Acids, bases, and salts, when dissolved in water, decompose into particles charged with electricity.*

The decomposition or dissociation of acids, bases, and salts when in solution is called *ionization*. The electrically charged particles formed by ionization are called *ions*. Each ion is a portion of a molecule. Two kinds of ions are present in every electrolytic solution, i. e., electro-positive ions, or *cations*, and electronegative ions, or *anions*.

Ions, although formed by the dissociation of molecules, must not be confused with atoms. *Ions are electrically charged atoms or radicals*. The electric charge changes the atom or radical, so that the properties of ions are quite different from that of atoms. It must be understood that the electric charges on the ions in a solution do not come from the electricity that may subsequently be passed through the solution. The ions are electrically charged just as soon as the molecules decompose in the solution.

According to the theory of ionization, an *acid* is a compound whose solution contains hydrogen ions ( $H^+$ ), while a *base* is a compound whose solution contains hydroxyl ions ( $OH^-$ ). Salts may be defined in several ways. For example, they are compounds (other than water) resulting from neutralization of acids and bases. Again, they are formed

by substituting (1) a metal for the hydrogen of an acid or (2) a nonmetal for the hydroxyl of a base. *According to the theory of ionization*, salts may be defined in two ways. *First*, salts are electrolytes which in solution yield neither hydrogen nor hydroxyl ions. *Second*, salts are compounds finally formed by the union of the positive ion of a base and the negative ion of an acid.

Neutralization, we learned, is a chemical change in which an acid and a base interact and form a salt and water. *Neutralization*, interpreted by the theory of ionization, is the combining of hydrogen and hydroxyl ions to form molecules of water.

Water itself conducts electricity very slightly indeed. If solutions are subjected to the action of an electric current, the results vary. Solutions of some substances do not conduct electricity at all, whereas solutions of certain substances conduct electricity readily; these substances are acids, bases, and salts. Thus we see that only solutions of *acids, bases, and salts are electrolytes*. The fact is readily interpreted by the theory of ionization. Solutions of acids, bases, and salts, according to the theory, contain electrically charged particles—ions; whereas solutions of other substances do not. Hence, when an electric current is introduced into solutions of acids, bases, and salts, particles *already charged with electricity* are there to conduct the current. No such particles are in solutions of other substances, hence such solutions cannot conduct an electric current.

*Electrolysis* is the term applied to the series of changes accompanying the passage of an electric current through a solution of an acid, base, or salt. Or it may be described as ionic migration induced by the electric current, the ions moving to their respective electrodes, where they are transformed into atoms or radicals which escape wholly or in part as elements, or which form various products by interaction with the water of the solution.

Except in unusual cases solutions of acids, bases, and salts contain both molecules and ions. That is, not all the molecules dissociate into ions. The degree of dissociation depends upon the concentration of solution and also upon the electrolyte itself. In concentrated solution the number of molecules is large. As the solution is diluted, more and more molecules dissociate into ions. If a solution of a substance is *strong*—it contains a large percentage of ions; if *weak* it contains a small percentage of ions.

## QUESTIONS FOR STUDY

1. State the theory of ionization. Illustrate it. State briefly the facts that support the theory.
2. Distinguish between an ion and an atom of sodium. How is each represented?
3. Interpret neutralization by the theory of ionization. Write the fundamental equation for neutralization.
4. Define acid, base, and salt in terms of the theory of ionization.

**Experiment 22. Alkalies.**

**Apparatus.** Six test tubes; two beakers; evaporating dish; ring stand with ring; wire gauze; Bunsen burner; glass rod.

**Material.** Washing soda; borax; sodium sulphate; baking soda; solution of sodium hydroxide, 10 g. to 100 c. c.; solution of ammonium hydroxide, 1 to 3; pieces of cloth with small grease spots; cotton cloth; woolen cloth; red litmus paper.

*a. Alkaline reactions.*

Dissolve a little of each of the following substances in water, and test the action of its solution on litmus, washing soda,  $\text{Na}_2\text{CO}_3$ ; borax,  $\text{Na}_2\text{B}_4\text{O}_7$ ; sodium sulphate,  $\text{Na}_2\text{SO}_4$ ; baking soda,  $\text{NaHCO}_3$ .

*Results in each case?*

*Are any of these substances bases?*

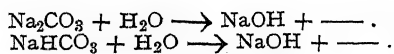
*Give reasons for your statement.*

*Name those which are alkalies.*

*Why are they alkalies?*

Alkalies which are not bases produce their effect on litmus because, on being dissolved, they react in a slight degree with water and form a very small quantity of base. The acid that is produced at the same time is relatively much weaker than the base.

*Complete the equations:*

*ib. Alkalies as solvents for grease.*

Put into a beaker a piece of cotton cloth, on which a very small grease spot has been made. Add sodium hydroxide solution, and boil the contents for several minutes. Remove the cloth and examine it to see if the grease spot has been affected.



*Result?*

Using a solution of borax as the alkali, repeat the experiment.

*Result?*

Which of the two alkalies has the greater grease-dissolving power?

*c. Action of alkalies on cotton and woolen goods.*

Put a small piece of each kind of goods into separate test tubes, add sodium hydroxide solution, and boil the contents for several minutes.

*Results?*

*Which kind of goods is changed more by the strong alkali?*

*d. Testing alkalies for volatility.*

A volatile substance is one which turns completely into a gas at ordinary temperatures, or with slight heating. Put 5 to 10 drops of sodium hydroxide solution into an evaporating dish, and heat until no liquid remains.

*Is there a residue?*

*Is sodium hydroxide volatile?*

In a similar way, heat a few drops of a solution of washing soda.

*Is this alkali volatile?*

By a third test, determine whether a solution of ammonium hydroxide is volatile.

*Result?*

*If these three alkalies were applied to clothing, which would evaporate, and which would remain on the cloth?*

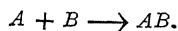
*Which of them would be most desirable to use in cleaning a grease spot or removing an acid stain from clothing?*

*Give reasons for your answer.*

## CHAPTER XVIII

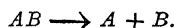
### CHEMICAL EQUILIBRIUM

We have regarded chemical reactions thus far as taking place only in one direction. Two substances,  $A$  and  $B$ , unite to form the compound  $AB$ , and we have written the equation expressing the reaction in the following manner:



This regards the compound  $AB$  as the static or unchanging condition into which the elements  $A$  and  $B$  have passed when they combine.

This is frequently not the whole truth. The compound  $AB$  often undergoes decomposition at the same time that it is being formed, giving again the elements  $A$  and  $B$ . This reaction would be represented as follows:



The reaction which took place originally between the elements  $A$  and  $B$  is exactly reversed in the second reaction. Such reactions, of which there is an unlimited number, are known as *reversible reactions*. Indeed, some are of the opinion that all chemical reactions are reversible, the original reaction in some cases proceeding, however, very rapidly, while the reverse action proceeds very slowly. This gives us the key to the formation of a certain amount of the substance  $AB$  from the elements  $A$  and  $B$  when the reaction is reversible. At first we have only the elements  $A$  and  $B$ . These begin to combine and form the compound  $AB$  with a certain velocity which at first is very great, but which becomes gradually slower and slower as the amounts of  $A$  and  $B$  become less and less. At first there is none of the compound  $AB$  present,—only the uncombined elements. When  $AB$  begins to form, the reverse action resulting in its decomposition into  $A$  and  $B$  begins, but at first has very small velocity. As the amount of the compound  $AB$  increases, the velocity with which it is decomposed also increases. The result is that the velocity of the original reaction is

becoming less and less, while the velocity of the reversed reaction is becoming greater and greater. After a time the two velocities become equal and we have then the condition described as *equilibrium*.

When equilibrium is reached, it does not mean that the reaction has ceased or that the reversed reaction has ceased, but that the two are taking place with the same velocity,—just as much of the compound *AB* *decomposing* in a given time as is *formed* in a given time. This is the same as to say that the condition of equilibrium is not a static condition as was for a long time supposed, but is a *dynamic condition*.

Let us apply this idea to the following reaction: Hydrogen and sulphur combine forming hydrogen sulphide, with a velocity which becomes less as the quantity of the elements present decreases. Hydrogen sulphide decomposes into hydrogen and sulphur with a velocity which becomes greater as the amount of hydrogen sulphide present increases; after a time just as much hydrogen sulphide is being decomposed in a given unit of time as is formed in the same time. Equilibrium between the two reactions is then established.

When equilibrium is established we have the maximum amount of hydrogen sulphide formed, which, under the conditions, could ever be formed.

All chemical reactions in which the yield is less than one hundred per cent are reversible, and since this theoretical yield is never quite fully realized, all reactions are probably, strictly speaking, reversible.

In some cases, however, the combination is so nearly complete that we must regard the velocity in one direction as infinitely great with respect to the velocity in the reverse direction. In such cases we would have, when equilibrium was reached, nearly all of *A* and *B* combined to form the compound *AB*, while a very slight amount of *AB* was decomposed into *A* and *B*. This is the condition in most reactions where a solid is precipitated. The solid is formed with a velocity which is far too great to measure, and the reaction proceeds nearly to the end before equilibrium between the two opposite reactions is established.

## QUESTIONS FOR STUDY

1. What do we mean by a reaction being in equilibrium? in dynamic equilibrium?
2. What is a reversible reaction? What is the proportion of the products to reacting substance?
3. Write two equations to show the difference between a reversible reaction and a reaction which goes to an end.

**Experiment 23. Neutralization.**

**Apparatus.** Two test tubes; glass stirring rod; porcelain evaporating dish; wire gauze; ring stand with one ring; Bunsen burner; glass plate; dropper.

**Materials.** Sodium hydroxide; blue and red litmus paper; hydrochloric acid, 1 to 10; potassium hydroxide; nitric acid, 1 to 10.

a. Put a small piece of sodium hydroxide (about half a gram) into a test tube half filled with water. As soon as the sodium hydroxide has dissolved, fill the tube with water and thoroughly mix the solution by pouring it back and forth from one test tube to another several times. Now pour half of the solution down a glass stirring rod into an evaporating dish.

Lay a piece of blue litmus paper and a piece of red litmus paper on a glass plate. Touch each kind of paper with the end of the stirring rod wet with the solution.

*What change do you observe?*

*Solutions producing such a change are called ———.*

Wet the end of a glass stirring rod with dilute hydrochloric acid and touch each kind of litmus paper with the wet end of the stirring rod.

*Result?*

*This reaction is characteristic of all water solutions of ———.*

Add the acid with the medicine dropper, a few drops at a time to the solution of sodium hydroxide in the evaporating dish. Stir the liquid thoroughly after each addition of acid, and, after stirring, touch a piece of blue litmus paper with the wet end of the stirring rod. Continue the addition of acid until you observe a change in the color of the litmus paper.

*The liquid now gives an ——— reaction.*

Pour a little of the sodium hydroxide solution remaining in the test tube into the evaporating dish, stir the liquid thoroughly, and touch a piece of each kind of litmus paper with the wet end of the stirring rod. If no change in the color of the red litmus paper takes place, add a little more of the sodium hydroxide solution.

*This liquid now gives an ——— reaction.*

Now add the acid, a drop at a time, until the resulting liquid changes the color of neither blue nor red litmus paper. The solution is now neutral.

Evaporate the liquid by setting the evaporating dish on a wire gauze over the flame of a Bunsen burner.

*What is the color of the residue?*

Taste the residue.

*What is it?*

*Complete the equation:*



b. Dissolve about half a gram of potassium hydroxide in a test tube full of water, and proceeding as in *a*, neutralize the solution with dilute nitric acid.

Evaporate the neutral solution to dryness. The residue is potassium nitrate,  $\text{KNO}_3$ .

*Write the equation for the neutralization.*

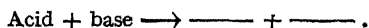
Potassium nitrate is called a salt, that is, it is a compound formed by the combination of a metal with an acid radical (an acid minus its replaceable hydrogen).

*What radical is present in every base?*

*What element is contained in every acid?*

*What become of the characteristic radical of a base and the characteristic element of an acid during neutralization?*

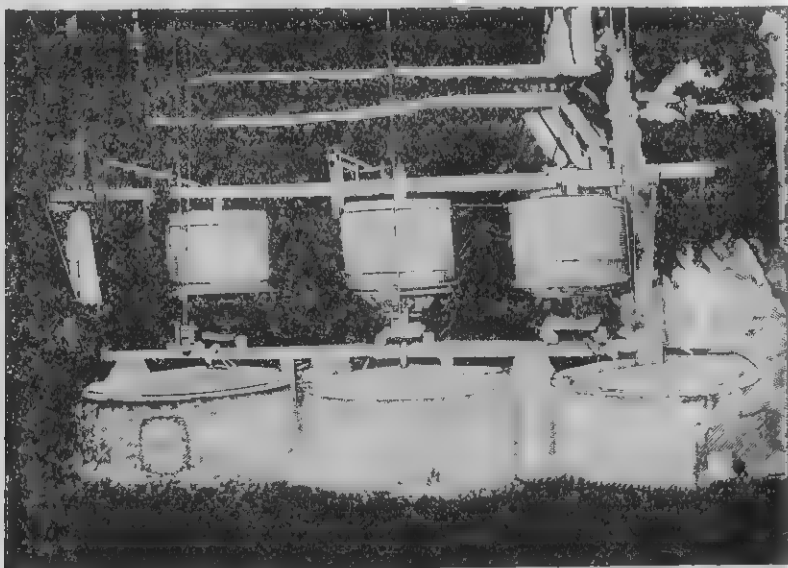
*Complete the general statement concerning neutralization:*



## CHAPTER XIX.

### SODIUM AND POTASSIUM COMPOUNDS

*Sodium chloride* ( $\text{NaCl}$ ) is the most important compound of sodium. It is familiar under the name of *salt* or *common salt*. The presence of salt in the ocean, in lakes and springs, and in the earth is mentioned in the oldest historical records. It is one of the most abundant sub-



From *China Journal of Science and Arts*

Fig. 31. Salt Making in West China

Behind and above the three gas-heated boiling pans are reservoirs for the brine waiting to be run into the pans.

stances, and is the chief of sodium compounds. Salt is an essential ingredient of the food of men and animals. Besides its universal domestic use, enormous quantities are used in making sodium carbonate, sodium hydroxide, and hydrochloric acid.

*Sodium carbonate* ( $\text{Na}_2\text{CO}_3$ ) is next to sodium chloride in importance. Formerly it was obtained from the ashes of marine plants, hence the old name *soda ash*; sodium chloride is now the source. Besides being

widely used as a cleansing agent, enormous quantities are consumed in the manufacture of glass, soap, sodium hydroxide, and many other useful substances.

*Sodium bicarbonate* ( $\text{NaHCO}_3$ ) is prepared by the Solvay process, or by treating sodium carbonate solution with carbon dioxide. It is an essential ingredient of baking powders used to raise bread, cake, and other food.

*Sodium hydroxide* ( $\text{NaOH}$ ), or caustic soda, is a white, crystalline, brittle, corrosive solid. Its solution is strongly alkaline and disinte-



From *China Journal of Science and Arts*

Fig. 32. There are five lines of bamboo pipes which conduct the brine to the boiling pans.

grates many substances, hence the name *caustic*. It is a strong base and its solution contains a high percentage of hydroxyl ions. Immense quantities are used in making soap, paper pulp, phenol, dyestuffs, and in refining petroleum.

*Sodium tetraborate* ( $\text{Na}_2\text{B}_4\text{O}_7$ ), or *borax*, is used chiefly in the manufacture of enamels for coating ironware. Some is used for preserving meat, fish, cheese, and other foods because it prevents the growth of certain bacteria. A solution of borax has a slightly alkaline

reaction owing to hydrolysis; hence it is sometimes used instead of soap as a cleansing agent. Considerable borax is consumed as an ingredient of ointments, lotions, and toilet powders.

*Sodium nitrate* ( $\text{NaNO}_3$ ) is found abundantly in Chile and is often called *Chile saltpeter*. Large quantities are used as a fertilizer, either alone or mixed with compounds of potassium or phosphorus.

Potassium compounds are found in many rocks and minerals, in wood ashes, and in the deposits in wine casks. Sea water and mineral waters



From *China Journal of Science and Arts*

Fig. 33. Weighing in Crystal Salt for the Retail Market

contain potassium salts, particularly potassium chloride and potassium sulphate. Extensive beds of potassium salts are found in Germany, especially at Stassfurt.

Potassium chloride ( $\text{KCl}$ ) resembles sodium chloride, and is used chiefly to prepare other potassium salts and as an ingredient of fertilizers.

Potassium nitrate ( $\text{KNO}_3$ ) also called *niter* and *saltpeter*, is formed in the soil of many warm countries by the decomposition of nitrogenous organic matter.



At a high temperature, potassium nitrate gives up oxygen readily, especially to charcoal, sulphur, and organic matter. This oxidizing power leads to its extensive use in making gunpowder, fireworks, matches, explosives, and in many chemical operations.

*Potassium chlorate* ( $\text{KClO}_3$ ) is used to prepare oxygen and in the manufacture of matches and fireworks.

*Potassium carbonate* ( $\text{K}_2\text{CO}_3$ ) is very soluble in water, and the solution, like that of sodium carbonate, has a strong alkaline reaction, owing to hydrolysis. It is used extensively in the manufacture of hard glass, soft soap, caustic potash (potassium hydroxide), and other potassium compounds.

*Potassium hydroxide* ( $\text{KOH}$ ) resembles sodium hydroxide in properties. Like sodium hydroxide, it dissolves readily in water with evolution of heat, forming a strongly alkaline solution. Its solutions corrode and disintegrate animal and vegetable matter and many mineral substances, hence the name *caustic potash*.

*Potassium cyanide* ( $\text{KCN}$ ) and *potassium sulphate* ( $\text{K}_2\text{SO}_4$ ) resemble the corresponding sodium compounds.

Potassium, like nitrogen and phosphorus, is essential to the life of plants and animals. The ash of many common grains, vegetables, and fruits contains potassium carbonate, which is formed from the complete organic potassium compounds in them. Potassium salts are taken from the soil by plants and must be returned if the soil is to be productive. Sometimes wood ashes, or the sulphate and chloride, are applied to the soil. Usually the potassium salts are supplied in the form of fertilizer. Experiments show that many soils need potassium salts as plant food.

#### QUESTIONS FOR STUDY

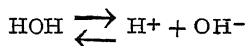
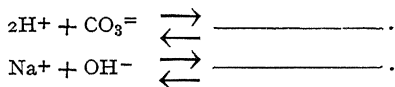
1. What is (a) soda? (b) soda ash? (c) sodium carbonate? (d) washing soda? (e) "alkali"? (f) acid sodium carbonate? (g) baking powder? (h) baking soda? (i) caustic soda?
2. Why does sodium carbonate form an alkaline solution?
3. Give the name and formula of the ions in dilute solutions of (a) sodium hydroxide, (b) potassium nitrate, (c) sodium chloride, and (d) potassium sulphate.

**Experiment 24. Salts that Are Not Neutral.****Apparatus.** Two test tubes.**Material.** Sodium carbonate; copper sulphate; ferric chloride; dilute solutions of aluminium sulphate, potassium chloride, borax, potassium nitrate, zinc sulphate, and ammonium sulphide; red and blue litmus papers.

a. Fill the curved portion of a test tube with powdered sodium carbonate,  $\text{Na}_2\text{CO}_3$ . Add water till the tube is two-thirds full and shake it until the sodium carbonate is dissolved. Into the solution dip a strip of blue litmus and a strip of red litmus.

*Record the result in the table.*

Water is *very slightly* dissociated into its ions according to the equation:

*Write the equation for the dissociation of sodium carbonate.**What four kinds of ions are present in the solution?**Complete the following:*

Carbonic acid has only a slight tendency to ionize, while sodium hydroxide ionizes in much greater degree. Therefore the first of these two actions tends to remove from the solution the  $\text{H}^+$  ions of the water in greater degree than the second tends to remove the  $\text{OH}^-$  ions.

*Which of these two ions tends to remain in excess?**How does this explain the action of sodium carbonate solution on litmus?*

The practical effect of dissolving sodium carbonate in water is to form, in slight degree, undissociated carbonic acid and dissociated sodium hydroxide.

Sodium carbonate is formed by neutralizing a \_\_\_\_\_ acid with a \_\_\_\_\_ base. Such a salt gives an \_\_\_\_\_ reaction in water solution.

Similarly test a solution of sodium tetraborate (borax) with the litmus papers.

*Record the results in your table.*

*Why would you expect sodium tetraborate to give such a reaction?*

b. Test a solution of copper sulphate and a solution of ferric chloride,  $\text{FeCl}_3$ , with red and blue litmus paper.

*Record the results in your table.*

*In each case, name the base formed that has little tendency to dissociate again.*

*What ion is responsible for the litmus reaction obtained?*

*Complete the statement:*

Salts with an acid reaction are formed by neutralizing a ——— acid with a ——— base.

c. Test solutions of aluminium sulphate, potassium chloride, potassium nitrate, zinc sulphate, and ammonium sulphide by red and blue litmus. Record the results in the table.

Aluminium hydroxide and zinc hydroxide are weak bases. Hydro-sulphuric acid,  $\text{H}_2\text{S}$ , and tetraboric acid are weak acids.

*Complete:*

weak acid + strong base  $\longrightarrow$  salt with ——— reaction.

strong acid + weak base  $\longrightarrow$  salt with ——— reaction.

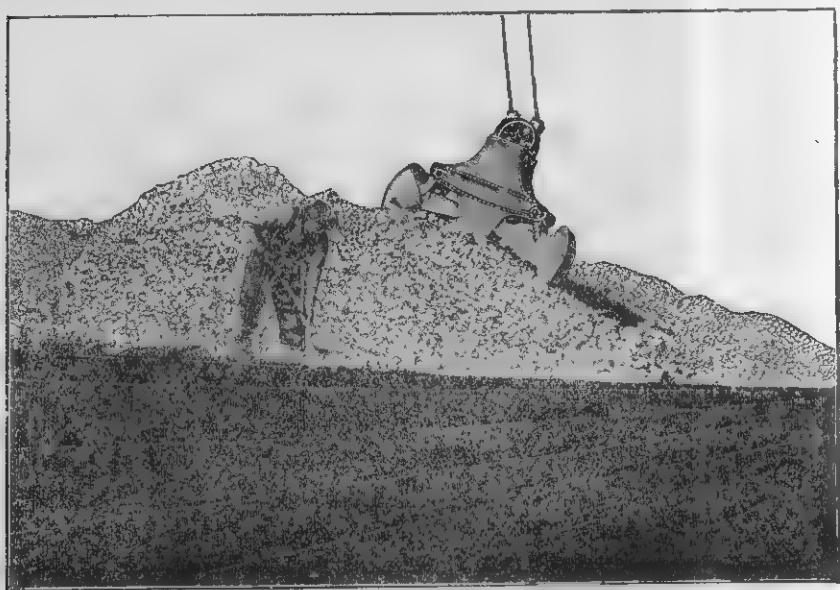
strong acid + strong base  $\longrightarrow$  salt with ——— reaction.

SALT	ACTION WITH BLUE LITMUS	ACTION WITH RED LITMUS
Sodium carbonate		
Copper sulphate		
Ferric chloride		
Aluminium sulphate		
Potassium chloride		
Sodium tetraborate		
Potassium nitrate		
Zinc sulphate		
Ammonium sulphide		

## CHAPTER XX

### SULPHUR AND SULPHIDES

The element sulphur has been known from very ancient times, having been used as a medicine by the Greeks and Romans. It is widely distributed in nature and occurs in large quantities in the uncombined



*From The National Geographic Magazine*

Fig. 34. MOST OF THE SULPHUR AMERICA USES COMES FROM TEXAS, U. S. A.

By hot water and steam, forced down holes as deep as 1,500 feet, sulphur is melted; then pumped into huge wooden vats at the mines, where it solidifies. For shipment, the sides of the vats are removed, and the sulphur is sometimes blasted to load into railway cars. Ships are loaded by steam shovels from the bins on the docks at Galveston, the chief sulphur port of the United States, which is near the Matagorda and Brazoria County deposits, two of the greatest sulphur-producing fields in the world. Last year America's production of sulphur exceeded 2,000,000 tons, of which Texas supplied practically all. Sulphur now is used in the manufacture of sulphuric acid, matches, gunpowder, and fertilizer.

form, especially in the neighborhood of volcanoes. Large deposits of free sulphur occur in Sicily and in the southern part of the United States. It exists in the pure state in many parts of West and North

Yünnan province of China as a deposit on the mouth of hot springs. It also occurs in Shansi, Shantung, Manchuria, and Kweichow.

Sulphur occurs in nature also as *calcium sulphate* ( $\text{CaSO}_4$ ), *gypsum* ( $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ ), *iron pyrite*, or *fool's gold* ( $\text{FeS}_2$ ), *galenite* ( $\text{PbS}$ ), and *zinc blende* ( $\text{ZnS}$ ). All plants and animals also contain sulphur. Without it they could not live. Sulphur is a constituent of albumen. Thus muscular tissues, horns, hoofs, hair, nerves, eggs, and seeds contain sulphur. In urine sulphur is present in the form of sulphates, for, as



From *Popular Mechanics* (Chicago)

Fig. 35. -THREE THOUSAND ACRES OF IT IN SALT LAKE, U. S. A.

Pulling a rope of crude bitumen from one of the wells; the material, practically pure, possesses high viscosity and great ductility, as the picture shows. Chemically, it is "saturated sulphur oil created by the decay of fossil remains, and sealed into clay beds of this region." Is used for making all kinds of rubber goods.

the animal lives, the sulphur compounds in its tissues are continually oxidized to sulphates, which are then eliminated by the kidneys. Many plants, as onions, cabbages, garlic, and mustard contain sulphur compounds. In order to thrive well they require a soil which contains an adequate amount of sulphur. When plants and animals decay, the sulphur compounds they contain gradually decompose, forming hydrogen sulphide when the sulphur is relatively abundant and access to the

air is limited, and sulphates when there is proper access of air. When these sulphates are again put into the soil, as in manuring or treating with gypsum, the rootlets of plants again take them up, and transform them into the various sulphur compounds of their tissues, and thus the *sulphur cycle* is completed.

Sulphur is used in the manufacture of gunpowder, lime-sulphur sprays for killing insects and fungous growths, carbon disulphide, sulphur dioxide, sulphuric acid, and in vulcanizing rubber.

It combines with the majority of the elements; and the sulphides can be prepared by direct union of the elements, just as the oxides can. The sulphides and oxides resemble each other in their chemical behavior. Both unite with hydrogen to form a compound containing two atoms of the latter element; that is, both are *divalent*.

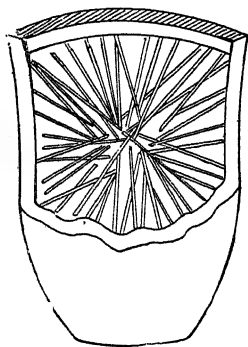


Fig. 37. The side of a crucible broken away to show needle-shaped, prismatic crystals of sulphur. Melts at  $119^{\circ}$  and has density of 1.96. At all temperatures below  $96^{\circ}$  the needle-shaped crystals break up more or less rapidly into crystals of rhombic sulphur.

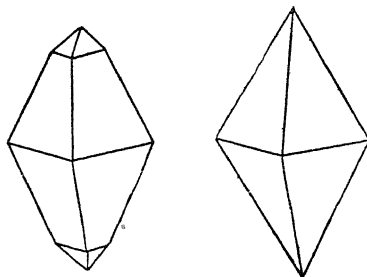


Fig. 36. Rhombic form of sulphur. Melts at  $114.5^{\circ}$ . Has density of 2.06.

The compound of hydrogen and sulphur, *hydrogen sulphide*,  $\text{H}_2\text{S}$ , is a gas formed in nature by the decay of organic matter containing sulphur—as we noticed before; it also occurs in volcanic gases and in the water of sulphur springs.

In the laboratory, hydrogen sulphide is prepared by the action of dilute hydrochloric acid on iron sulphide. It is a colorless gas, with a characteristic, unpleasant odor, fairly soluble in water, heavier than air. It burns in an abundance of air to form water and sulphur dioxide, but in a limited supply of air it forms water and sulphur.

The *water solution* is called hydrosulphuric acid; it is a weak acid and forms sulphides with most metallic solutions. It is used in chemical analysis to separate metals.

## QUESTIONS FOR STUDY

1. Describe how you would proceed to prove a yellow powder to be sulphur.
2. How does sulphur occur in nature? What are the uses of sulphur?
3. Discuss the rôle of sulphur in plants and animals.
4. Write the names and formulas of four metallic sulphides and of the corresponding oxides.
5. What is meant by saying that hydrosulphuric acid is weak?

**Experiment 25. Forms of Sulphur.**

**Apparatus.** Two test tubes; watch glass; test-tube holder; Bunsen burner; magnifying glass (if possible); pan of water.

**Material.** Roll or powdered sulphur; carbon disulphide; 4-inch filter paper.

*a. Caution! Carbon disulphide is a volatile liquid that takes fire easily. It should never be used near a flame.*

Pour 5 c. c. of carbon disulphide into a test tube. Add a piece of roll sulphur the size of a pea and shake the tube.

Pour the clear liquid into a watch glass and set it aside to evaporate in a part of the laboratory some distance from a flame. Using a magnifying glass, examine the crystals of sulphur.

*Make a drawing of a crystal having a symmetrical form.*

You have recrystallized roll sulphur under conditions that yield separate crystals. Sulphur that crystallizes as you have just observed is called *rhombic sulphur*, one of the four allotropic forms of sulphur.

*b.* Fold a piece of filter paper as you would to fit a funnel, and lay it aside for future use. Also have a dish of water ready for use

Half fill a test tube with small pieces of roll sulphur. Carefully melt the sulphur by holding the tube in an inclined position about four inches above a *small* flame. Rotate the tube slowly while the melting proceeds.



Fig. 38. Folding a Filter

*What is the color and consistency of the first portion of liquid obtained?*

This color should be retained during the melting of all the sulphur. The color should at no time be darker than a light amber.

Holding the filter by the edge, pour the melted sulphur into it. As soon as crystals have formed from the edge to the center of the surface,

pour into the water in the pan the part of the sulphur that is still in a melted condition. *Immediately* unfold the filter paper.

*Make a drawing of one of the more perfect crystals as it is seen under a microscope.*

This form of sulphur is known as *prismatic* sulphur.

Keep some of the crystals for a few days and then examine them.

*What changes do you observe?*

c. Half fill a test tube with small pieces of roll sulphur, and holding the test tube with a holder (which could be made of a folded strip of paper), raise the temperature of the sulphur until it commences to boil. Meanwhile, tip the tube slightly from time to time and note the important changes that take place in the color and consistency (degree of fluidity) of the sulphur.

*What changes in the color and consistency of the sulphur did you note from the time it melted until it commenced to boil?*

In the next operation the sulphur will probably take fire. *Do not be frightened.* The sulphur will burn quietly. Do not spill it on the desk. Pour the *boiling* sulphur slowly into cold water, keeping the mouth of the tube moving in a circle so that a thread of sulphur will form in the water. Examine the thread of sulphur. It is the plastic form of sulphur.

*What color is it?*

*Is it hard or soft? elastic or brittle?*

Keep it for several days and note any change in properties

### Experiment 26. Preparation of Metallic Sulphides.

**Apparatus.** Test tube; ring stand with one ring; wire gauze with asbestos center; Bunsen burner; iron clamp.

**Material.** Powdered sulphur; copper wire or sheet copper; clean and fine iron filings; zinc dust; hydrochloric acid, 1 to 4.

a. Fill the curved portion of a test tube with sulphur and heat it to boiling. Insert a strip of thin sheet copper (or fine copper wire) into the boiling sulphur.

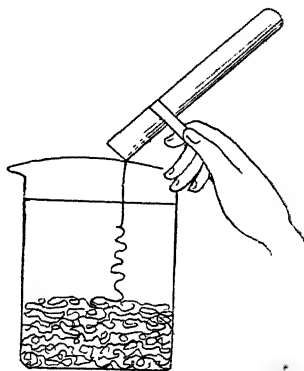


Fig. 39. Pouring melted sulphur into cold water makes plastic sulphur.



*Result?*

Withdraw the strip and compare its color, luster, and flexibility with the color, luster, and flexibility of copper.

*Result?*

*What is the name of the compound formed by the combination of copper with sulphur?*

*Write the equation representing the formation of this compound.*

b. Mix thoroughly one part by volume of finely powdered sulphur with two parts of fine iron filings. Put the mixture into a test tube and heat the lower end of the tube just sufficiently to color the flame yellow. When the contents of the tube commence to glow, withdraw the flame.

*Does the chemical action continue? Why do you think so?*

Break the tube and examine the contents.

*How does it differ from sulphur in appearance?*

*What gas is produced when hydrochloric acid reacts with iron?*

Add one drop of hydrochloric acid to the substance taken from the broken test tube. Cautiously smell of the gas.

*What evidence is there that the substance taken from the tube is not iron?*

*Write the equation for the reaction.*

c. This should be performed by the teacher. Mix thoroughly a pinch of sulphur with an equal bulk of powdered zinc. Placing the mixture in a conical pile on asbestos and holding the burner at arm's length, cautiously ignite the pile from above.

*Result? Write the equation for the reaction.*

Compare the action of copper and zinc with sulphur with the action of these metals with oxygen.

## CHAPTER XXI

### OXIDES AND ACIDS OF SULPHUR

Sulphur forms two well-known compounds with oxygen; sulphur dioxide ( $\text{SO}_2$ ), sometimes called *sulphurous anhydride*; and sulphur trioxide ( $\text{SO}_3$ ), frequently called *sulphuric anhydride*. It may be prepared in the laboratory by heating sulphur or certain compounds containing sulphur in air or oxygen, or by the reduction of hot concentrated sulphuric acid with certain metals such as copper.

Sulphur dioxide has a marked tendency to combine with other substances, and is therefore an active substance chemically. It can take oxygen away from some other substances, and is therefore a good reducing agent. It has a great tendency to combine with water, forming *sulphurous acid*,  $\text{H}_2\text{SO}_3$ , which has all the properties typical of a weak acid. A solution of sulphurous acid acts as a good reducing agent, as it has the power of taking up oxygen from the air or from substances rich in oxygen, being changed by this reaction into sulphuric acid. It is also used as a bleaching agent and disinfecting agent. The salts of sulphurous acid are called *sulphites*.

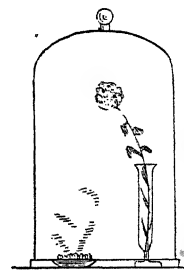
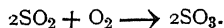


Fig. 40. Bleaching a Flower with Sulphur Dioxide

When sulphur dioxide and oxygen are heated together at a rather high temperature, a small amount of *sulphur trioxide*,  $\text{SO}_3$ , is formed, but the reaction is slow and incomplete. If, however, the heating takes place in the presence of very fine platinum dust, the reaction is rapid and nearly complete:



It has been found that many chemical reactions, such as the union of sulphur dioxide with oxygen, are much influenced by the presence of substances which do not themselves seem to take a part in the reaction and are left apparently unchanged after it has ceased. These reactions go on very slowly under ordinary circumstances, *but are greatly hastened*

by the presence of the foreign substance. Substances which *increase the speed of reactions* in this way are said to act as *catalytic agents*, or *catalyzers*, and the action is called *catalysis*. Just how the action is brought about is not well understood, but the part played by the catalyzer is no doubt different in different cases.

Enormous quantities of sulphuric acid are manufactured by two processes, known as the lead chamber process and the contact process.

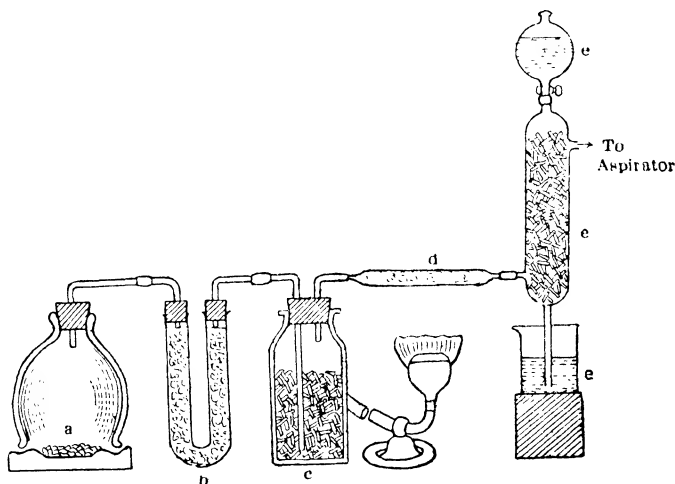


Fig. 41. Apparatus that can be used in the laboratory to show how sulphuric acid is manufactured on a large scale by the contact process.

*a.* Sulphur burns in air which is drawn into apparatus by aspirator.

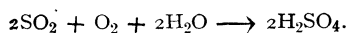
*b.*  $\text{SO}_2$  with excess of air passes into U-tube, filled with loose material to rid gas of sulphur dust.

*c.* Gases dried in bottle containing glass beads wet with concentrated  $\text{H}_2\text{SO}_4$ .

*d.*  $\text{SO}_2$  combines with  $\text{O}_2$  of excess air, aided by gently heated platinized asbestos.

*e.* Sulphur trioxide absorbed by concentrated  $\text{H}_2\text{SO}_4$ . Then water added.

In each process sulphur dioxide is oxidized to sulphur trioxide, which with water forms sulphuric acid. A general equation for the essential chemical change is:



In the lead chamber process (which is very cheap in operating costs, producing an acid of about 60%–70% purity) the oxidation is

accomplished by nitrogen oxides; and in the contact process it is hastened by a catalyst.

Sulphuric acid is an oily liquid, colorless when pure, but often brown from the presence of charred organic matter, such as dust and straw. It is nearly twice as heavy as water; and boils at  $338^{\circ}$ , decomposing to some extent and forming dense, white, suffocating fumes.

When sulphuric acid is mixed with water, much heat is evolved. *The acid should always be poured into the water and the mixture should be stirred*, otherwise the intense heat may crack the vessel or spatter the hot acid. This tendency to absorb water is shown in many ways. The concentrated acid absorbs moisture from the air and from gases passed through it. Organic substances, such as wood, paper, sugar, starch, and cotton are charred by sulphuric acid. Such compounds contain hydrogen and oxygen in the proportion to form water; these two elements are abstracted and carbon alone remains. Sulphuric acid also disintegrates the flesh, often causing serious burns, and if accidentally spilled on the hands or spattered on the face should be washed off immediately.

Sulphuric acid is one of the most important substances. Directly or indirectly it is used in hundreds of industries upon which the comfort, prosperity, and progress of mankind depend. On account of its varied and extensive uses, *sulphuric acid may be considered as one of the foundation stones of modern civilization*. It is used in the manufacture of many acids. It is essential in one process for the manufacture of sodium bicarbonate, which has many uses. Enormous quantities are consumed in making fertilizers, alum, and other sulphates, nitroglycerin, glucose, dyes, and in various parts of such fundamental industries as dyeing, bleaching, metal cleaning, refining, and metallurgy.

The salts of sulphuric acid are the *sulphates* (such as  $\text{Na}_2\text{SO}_4$ ), and the acid sulphates or bisulphates (such as  $\text{NaHSO}_4$ ). Most sulphates are soluble in water; only the sulphates of barium, strontium, and lead are insoluble, while calcium sulphate is but slightly soluble. Sulphates are used in many industries.

#### QUESTIONS FOR STUDY

1. Starting with sulphur, how would you prepare successively sulphur dioxide, sulphur trioxide, sulphuric acid, hydrogen sulphide, and sulphur?

2. What is an acid anhydride? Name two anhydrides of sulphur.
3. Mention two instances of catalysis.
4. Describe the properties of sulphuric acid. What use is made of this substance?

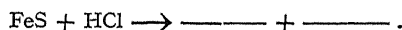
**Experiment 27. Preparation and Properties of Hydrogen Sulphide.**

**Apparatus.** Five test tubes; stopper and delivery tube to fit one of the test tubes; Bunsen burner.

**Material.** Iron sulphide; dilute hydrochloric acid; solutions of lead nitrate, cadmium nitrate, and hydrogen peroxide; litmus paper.

a. In a test tube provided with a stopper and a delivery tube place two or three small pieces of iron sulphide and cover with dilute hydrochloric acid.

*Complete the equation:*



Collect the gas by downward displacement in a dry test tube, remove the delivery tube, and light the gas in this test tube.

*What two products are formed when the gas burns in this way? (Examine the tube carefully for traces of moisture before it becomes heated by the flame of the burning gas.)*

*How do you know what these products are?*

*What does this show concerning the composition of the gas?*

b. If the gas in the test tube burned *quietly*, light the gas at the end of the delivery tube. *Cautiously* smell the gas around the flame.

*Are all the products formed the same as before?*

*State the reason for your answer.*

c. Place the end of the delivery tube in a test tube half filled with water and let the gas bubble through the water two or three minutes.

*How does the solution taste?*

*What effect has it on litmus?*

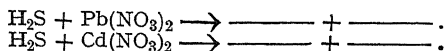
Using a small portion of the solution of hydrogen sulphide in each case, add a few drops of:

- (1) a solution of lead nitrate;
- (2) a solution of cadmium nitrate.

*Results?*

*Complete the equations:*

## OXIDES AND ACIDS OF SULPHUR



The cadmium compound precipitated is used as a pigment by artists. It is known as cadmium yellow.

*d.* Allow the gas from the generator to bubble through a solution of hydrogen peroxide until a decided effect is obtained.

*Result?*

*Complete the equation:*



### Experiment 28. Preparation and Properties of Sulphuric Acid.

**Apparatus.** Beaker; two test tubes; porcelain evaporating dish; flask, 50 c. c.; glass stirring rod; ring stand with ring and small clamp; gauze with asbestos center; Bunsen burner.

**Material.** Concentrated sulphuric acid; dilute hydrochloric acid; barium chloride solution, 1 to 20; sodium sulphate solution, 1 to 20; cane sugar; strips of zinc; fine copper wire or wire gauze; wood splinter.

*a. Action with water.*

*Caution! In mixing concentrated sulphuric acid with water, the acid should be slowly poured into the water, with constant stirring. The reverse method produces a dangerously explosive spattering.*

Pour a test-tubeful of water into a beaker. Into this water slowly pour one sixth of a test-tubeful of concentrated sulphuric acid, frequently stirring the mixture with a glass rod. Feel the outside of the beaker.

*What noticeable effect is produced?*

Keep for parts *b* and *d* the dilute sulphuric acid just prepared.

*b. Action with metals.*

Put a zinc strip into one sixth of a test-tubeful of concentrated sulphuric acid.

*Is there much action between the zinc and the concentrated sulphuric acid?*

Pour the contents of the tube into the sink and wash down the acid with water. Rinse off the zinc strip and return it to the test tube. Then pour upon it some of the dilute sulphuric acid made in part *a*.

*Describe the action.*

*Name the gaseous product, and write the equation for the reaction.*

*How does dilute sulphuric acid differ from the concentrated acid in its action with metals?*

Place some copper wire gauze or a small loosely rolled ball of fine copper wire in a small flask, and add a fifth of a test tube of concentrated sulphuric acid. Support the flask on an asbestos gauze on a ring stand (in a hood if possible, if not, be sure to open the windows of the laboratory) and loosely clamp the neck of the flask. Heat the flask carefully *with a small flame* until action commences. Then remove the flame.

*Describe the action.*

Cautiously smell the gaseous product.

*What is it?*

*What gas is usually liberated when an acid reacts with a metal?*

*Remembering that hot, concentrated sulphuric acid acts as an oxidizing agent, explain why we do not get this gas here.*

*From which of the original materials is the sulphur dioxide derived?*

*Name the salt formed.*

*Complete the equation*



Mercury and silver react similarly to copper with hot concentrated sulphuric acid.

*Write the equation for the reaction in the case of silver.*

c. *Dehydrating action.*

Pour one sixth of a test tube of sulphuric acid into an evaporating dish. Add a little more than enough cane sugar to soak up the acid. Allow the action to continue until a decided result is obtained.

*Describe the action that occurs.*

*Which element of the cane sugar molecules  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$  gives the color to the residue in the dish?*

*What elements did the concentrated sulphuric acid remove from the cane sugar molecules?*

Dip a wooden splinter into concentrated sulphuric acid.

*Results?*

*Explain the effect of the acid upon the wood, which is mainly cellulose, with the formula  $\text{C}_6\text{H}_{10}\text{O}_5$ .*

Complete the statement —

When sulphuric acid as a dehydrating agent on compounds. it removes from them — and — as —.

*Why is sulphuric acid used in drying gases?*

*d. Test for a sulphate.*

To a little sodium sulphate solution in a test tube, add a few c. c. of barium chloride solution.

*Describe the color and character of the compound produced.*

*Write the equation for its formation.*

Add dilute hydrochloric acid to the precipitated barium sulphate.

*Result?*

Using barium chloride solution and dilute hydrochloric acid, repeat the test with dilute sulphuric acid.

*Result?*

*State the test for a sulphate.*

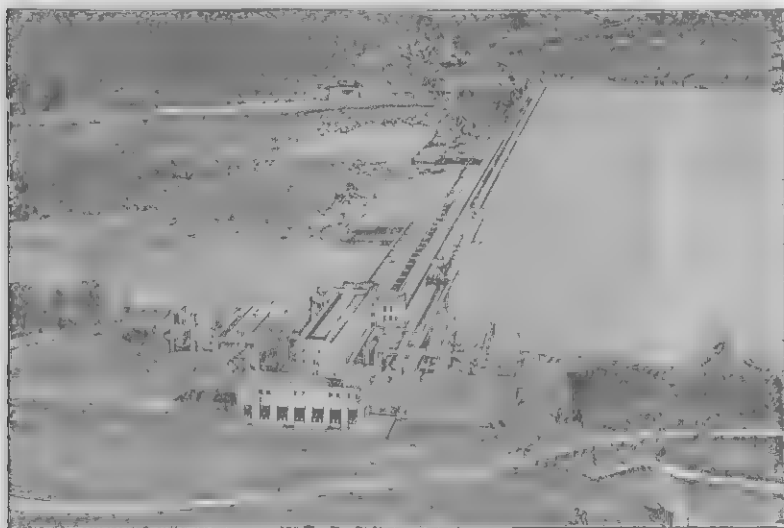


## CHAPTER XXII

### NITROGEN AND THE ATMOSPHERE

The *atmosphere*, or *air*, is a great mass of gas which envelops the earth and extends for several miles into space. We found that oxygen made up about one fifth of the air; practically all the remaining four fifths of the air is the elementary gas *nitrogen*. Nitrogen is found in many animal and vegetable substances essential to life, e. g., the compounds called proteins, which are indispensable ingredients of our food and also of the muscles and nerves of our bodies.

Nitrogen is a colorless gas, without taste or odor, and is a little lighter than air. It does not support combustion or sustain life.



From *The National Geographic Magazine*

Fig. 42. A Glimpse of the United States Government Nitrate Plant at Muscle Shoals for the Fixation of Nitrogen

Flames are extinguished by nitrogen and animals die in it, because the supply of oxygen is cut off. The fact that nitrogen quickly extinguishes a candle flame and kills a mouse was first observed in

1772 by a Scottish physician named Rutherford. The name *nitrogen* was given to it because it is a constituent of the important compound niter (i. e., saltpeter,  $\text{KNO}_3$ ).

Nitrogen is sometimes called an *inert element*, because it does not combine with elements at ordinary temperatures. At high temperatures and under special conditions, however, nitrogen forms many compounds. At high temperatures, or by the aid of electric sparks in the presence of a catalyst, nitrogen may be made to combine with other elements enabling us to convert the inert nitrogen of the air into compounds needed as fertilizers and explosives.



From *The Saturday Evening Post*

Fig. 43. The World's Largest Air Nitrogen Fixation Plant at Merseburg, Germany

Nitrogen, as well as oxygen, is vitally connected with life, though in a different way. But although we live in an atmosphere containing such a large proportion of this gas, we cannot assimilate it directly. The nitrogen we inhale (along with the oxygen) is exhaled again unused. The nitrogen needed by animals must be in the form of nitrogenous food existing in lean meat, fish, beans, wheat, and other grains.

Nor have plants, with few exceptions, power to obtain free nitrogen from the atmosphere. Most plants take up combined nitrogen from the soil in the form of nitrates or ammonia. Hence combined nitrogen is constantly being removed from the soil. In order to restore it, some nitrogen compound must be added, e. g., sodium nitrate,  $\text{NaNO}_3$ , calcium nitrate,  $\text{Ca}(\text{NO}_3)_2$ , ammonium chloride,  $\text{NH}_4\text{Cl}$ ,

ammonium sulphate,  $(\text{NH}_4)_2\text{SO}_4$ , or calcium cyanamide,  $\text{CaCN}_2$ : organic materials are often used, such as manure, dried blood, and meat or fish scraps. Any substance which restores nitrogen (or some other chemical element like phosphorus or potassium) to the soil is called a *fertilizer*.

Many experiments have shown, however, that leguminous plants such as peas, beans, and clover, take up nitrogen from the air by means of bacteria which are in nodules on their roots. Sometimes soil is treated with a preparation which contains nitrogen-forming bacteria.

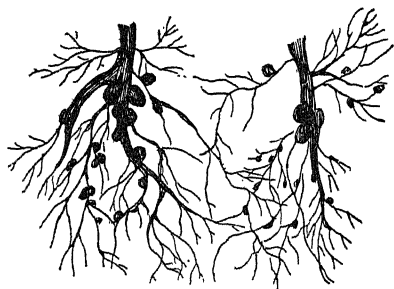


Fig. 44. Nodules of Nitrifying Bacteria on the Roots of Bean Plants

These may be easily observed by pulling up bean, pea, or clover plants.

Air is a mixture of several gases. Oxygen, nitrogen, and argon are the three ingredients that are always present in nearly constant proportions. Variable proportions of water vapor and carbon dioxide are always found, and also small quantities of compounds related to ammonia and nitric acid.

Near cities the air may contain considerable dust, sulphur compounds, and acids.

We know now that air is not a compound but a *mixture* of gases, for

1. The proportion of oxygen and nitrogen is not fixed, but varies from place to place. Therefore air does not have a constant composition and cannot be represented by a formula.

2. When nitrogen and oxygen are mixed in approximately the proportions that form air, the product is identical with air, but the act of mixing gives no evidence of chemical action, i. e., no heat is absorbed or given off.

3. When air is dissolved in water, a larger proportion of oxygen than nitrogen dissolves. If the oxygen and nitrogen were combined, the dissolved air would contain the same proportions of oxygen and nitrogen as air itself.

4. When air is liquefied and allowed to boil, the nitrogen boils off first. If air were a compound, liquid air would have a fixed boiling point.

## QUESTIONS FOR STUDY

1. What is the relation of nitrogen to the life of (a) animals? and (b) plants?
2. Give several proofs that air is a mixture.
3. When oxygen and nitrogen are mixed in the proportion in which they exist in the atmosphere, heat is neither given off nor taken on in the process. What important point does this suggest?
4. Can you suggest any reason why the growth of clover in a field improves the soil?

**Experiment 29. Preparation and Properties of Ammonia.**

**Apparatus.** Ring stand; clamp; test-tube rack; three test tubes, one fitted with one-holed stopper and delivery tube; beaker of water.

**Material.** Slaked lime; ammonium chloride; ammonium sulphate; sodium hydroxide solution; red and blue litmus paper.

*a. Preparation of ammonia.*

Take a little ammonium chloride in one hand and in the other a little slaked lime (dry). Smell of each. Rub the two together between the palms of the hands. Smell the mixture cautiously. *What does it smell like?* Bring a moist strip of litmus paper near the mixture?

*Result?*

*b.* Repeat, using ammonium sulphate and slaked lime. As before, smell, and apply the litmus test.

*Results?*

*c.* To a little sodium hydroxide solution in a test tube, add a small amount of ammonium chloride. Heat gently, smell cautiously, and test with litmus as before.

*Results?*

Ammonium chloride,  $\text{NH}_4\text{Cl}$ , and ammonium sulphate,  $(\text{NH}_4)_2\text{SO}_4$ , are salts.

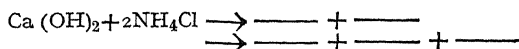
*Regarding the reaction between sodium hydroxide and ammonium chloride as a double replacement, name the two products formed.*

*Write the equation showing these two as products.*

Ammonia gas,  $\text{NH}_3$ , results from the decomposition of one of the products.

*Write the equation for this decomposition.*

Complete the equation for the action in part *a*.



*Write a single equation for the reaction in part *b*.*

*How could you prove that a substance given you was an ammonium salt?*

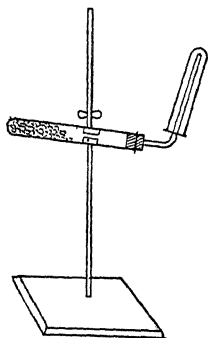


Fig. 45. Collection of Ammonia by Upward Displacement

*d.* On a piece of paper, mix a quarter of a test-tubeful of ammonium chloride with a quarter of a test-tubeful of slaked lime. Put the mixture into a test tube provided with a delivery tube for the collection of the gas by upward displacement. Warm the mixture very gently. Fill a dry test tube with the gas.

*e. Properties of ammonia.*

*What is the color of the ammonia gas?*

*Is this gas lighter or heavier than air?*

*f.* Place a test tube of ammonia mouth downward in a beaker of water.

*Result?*

*What does this show?*

### Experiment 30. Ammonium Compounds.

**Apparatus.** Test tube with stopper and single-bend delivery tube; four other test tubes; ring stand with one clamp; Bunsen burner.

**Material.** Ammonium chloride; slaked lime; concentrated hydrochloric acid; concentrated sulphuric acid; sodium hydroxide solution, 1 to 10; red litmus paper; blue litmus paper; labels.

On a piece of paper mix thoroughly 5 grams of ammonium chloride and 10 grams of dry slaked lime. Notice the characteristic odor of ammonia. Place the mixture in a dry test tube provided with a stopper and a delivery tube. Clamp it in a nearly horizontal position with the delivery tube pointing down. Have at hand 3 test tubes: (*a*) a dry test tube to which 2 drops of concentrated hydrochloric acid have been added; (*b*) a test tube containing 1 drop of sulphuric acid; (*c*) a test tube containing water to the depth of 1 inch.

a. Warm gently the test tube containing the mixture and bring the delivery tube into the test tube containing the drops of hydrochloric acid, until a solid is formed. This solid is ammonium chloride,  $\text{NH}_4\text{Cl}$ .

*Write the equation for the reaction.*

*Was there any heat developed when it formed?*

Label and reserve the material.

b. Similarly use the test tube containing the single drop of sulphuric acid. The solid formed in this case is ammonium sulphate,  $(\text{NH}_4)_2\text{SO}_4$ .

*Write the equation for the reaction.*

*Was there any heat developed when it formed?*

Label and reserve the material.

c. In the same way, bring the delivery tube into the test tube containing the water, but be careful that the delivery tube *does not touch the water*. Observe the top of the water.

*Is there any sign of action?*

*Is there any heat developed?*

Test the liquid with litmus. *Result?*

*Has the liquid any odor?*

The liquid is a solution of ammonium hydroxide,  $\text{NH}_4\text{OH}$ , an unstable base which is commonly used.

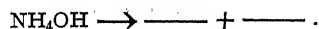
*Write the equation for its formation.*

d. Using half of this solution of ammonium hydroxide, boil the liquid for three minutes, while holding a piece of moistened litmus paper across, but not touching the mouth of the test tube.

*How do you know that something besides water vapor is escaping?*

Label and reserve the material. Ammonium hydroxide is very unstable. It decomposes, forming water and gaseous ammonia,  $\text{NH}_3$ .

Complete the equation:



e. Using the other part of the ammonium hydroxide solution obtained in part c, add to it 2 drops of hydrochloric acid, mix well,

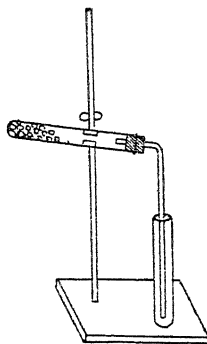


Fig. 46

and if there is any odor, warm the solution gently until the odor has disappeared. Label and reserve this solution for part *h*.

*Write the equation.*

*f.* Using the test tube containing the solid ammonium chloride (part *a*), warm it gently at the spot where the most solid seems to be.

*Result?*

*Would this have happened to sodium chloride?*

When the test tube has cooled, add a little water.

*Does it dissolve in water?*

*How does it react with red and blue litmus?*

Label and reserve the material for part *h*.

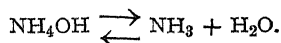
Ammonium compounds resemble chemically the compounds of potassium and of sodium. In the ammonium compounds, the group of atoms ( $\text{NH}_4$ ) acts like an atom of sodium or potassium.

*In this experiment, in what two cases has the ammonium hydroxide acted in the same manner as sodium hydroxide would have done?*

*In which case did it act differently?*

*h. Test for ammonium compounds.*

To detect ammonium compounds, convert the compound into the hydroxide and identify the unstable ammonium hydroxide by the ammonia gas resulting from its decomposition.



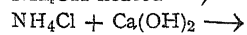
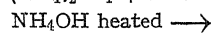
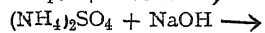
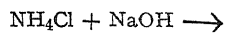
To each of the 4 materials reserved, after noticing whether there is any odor, add 5 or 6 drops of sodium hydroxide solution. Notice whether there is an odor, and, while holding a piece of moistened litmus paper just above the mouth of the test tube, warm the tube gently. Observe carefully the result in each case.

Tabulate results as shown below:

TABLE

MATERIAL IN TEST TUBE FROM	ODOR BEFORE ADDING NaOH	ODOR AFTER ADDING NaOH	ACTION WITH LITMUS
<i>d</i>			
<i>e</i>			
<i>f</i>			
<i>g</i>			

*Complete the equations:*



*Having performed the test for ammonia on these four samples, how could you now tell whether the original material was a chloride or a sulphate?*



## CHAPTER XXIII

### NITROGEN COMPOUNDS

The two most important common compounds of nitrogen are nitric acid and ammonia. Nitric acid may be made by subjecting a mixture of nitrogen and oxygen to the very high temperature of the electric arc and then treating with air and nitric acid. It may also be made from saltpeter, i. e., potassium nitrate, by treatment with strong sulphuric acid and heating. Instead of the somewhat costly potassium nitrate, sodium nitrate may be used. Thus nitric acid and sodium sulphate are produced instead of nitric acid and potassium sulphate.

Pure *nitric acid* is a colorless liquid having a pungent odor. In the sunlight the liquid turns yellowish in color due to partial decomposition. *Nitric acid is a very powerful acid and also a strong oxidizing agent.* It turns blue litmus red, discolors indigo solution, disintegrates cloth,

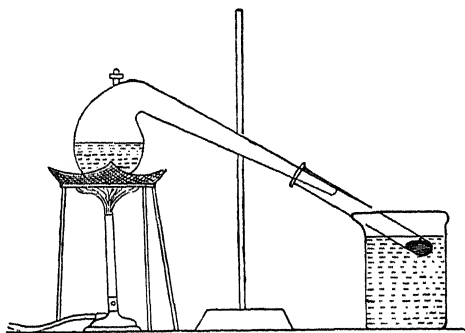


Fig. 47. The Preparation of Nitric Acid in the Laboratory

corrodes the flesh, and turns the skin yellow. Metals, with the exception of gold and platinum, are acted upon by nitric acid, being converted into either nitrates or oxides. The various salts of nitric acid are all called *nitrates*. The nitrates, especially those of potassium, ammonium, calcium, and sodium, are very important as fertilizers. While nitric acid does not occur in

nature as such, its salts, that is, the nitrates, are quite widely distributed in the soil. In the air nitrate of ammonium is also present. *Nitrates are exceedingly important as soil constituents*, for it is from them that plants get their supply of nitrogen. The nitrates in the soil are contained in the soil water which holds them in solution.

Nitric acid is also used in manufacturing sulphuric acid, in which case it serves to oxidize sulphur to its highest stage of oxidation.

Furthermore, in making collodion, guncotton, dynamite, and smokeless powder, nitric acid is used in large quantities.

*Nitrous acid* may be obtained from nitric acid by robbing the latter of a part of its oxygen. The salts of nitrous acid are called *nitrites*; they occur in the soil as a result of the decomposition of plant and animal matter. *All plants and animals contain nitrogen in combination*

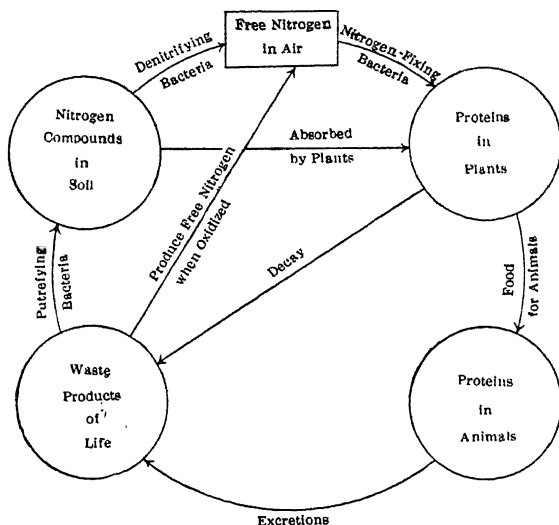


Fig. 48. Nitrogen Cycle

with carbon, hydrogen, oxygen, and minor amounts of sulphur and phosphorus, as we mentioned before. When the plant and animal tissues decay, nitrogen is set free in part, and in part is converted into ammonia. The ammonia is gradually oxidized to nitrites and then to nitrates. The latter represents the highest oxidation products. While the nitrates in the soil are a very important source from which plants derive their supply of nitrogen, they are by no means the only source.

Commercial ammonia and the ammonium compounds of commerce are in large quantity obtained as a by-product of the manufacture of illuminating gas from coal. In this process the coal, which represents the remains of the vegetation of the carboniferous age and other early geological periods, is heated out of contact with the air (called *dry* or *destructive distillation*) and a variety of gaseous products are formed, among which is ammonia.

Ammonia can now be produced by direct combination of hydrogen and nitrogen at high temperature and with the aid of catalysts.

Ammonia gas is colorless and only 0.59 times as heavy as air. It has a very characteristic, sharp penetrating odor, irritates the mucous membranes, and causes a flow of tears. One cannot inhale the gas. Animals soon die in it. The gas does not support combustion and does not burn in the air, but may be burned in oxygen. About 700 volumes of ammonia gas will dissolve in 1 volume of water at ordinary temperatures.

Ammonia water is used in the household in cleansing clothes and polishing metals. It does not attack the latter as much as acids do, and consequently is less objectionable.

Ammonia is formed wherever nitrogenous plant and animal remains are decaying. So, for instance, it is found in stables, in manure piles, in the leachings from the latter, and in the soil. When thus formed from rotting organic matter, ammonia at once combines with any acids that may be present, especially with carbonic acid gas, which is always present, being a constituent of the air. In the soil, then, ammonia is present in the form of ammonium salts; among these, ammonium nitrate, ammonium nitrite, ammonium carbonate, ammonium chloride, and ammonium sulphate are the most important. From ammonium salts dissolved in the waters of the soil, plants derive a considerable share of their supply of nitrogen, hence the value of ammonium salts as fertilizers.

#### QUESTIONS FOR STUDY

1. How may nitric acid be formed?
2. Of what use are nitrates in the soil?
3. What is ammonia? How is it prepared?
4. Why are ammonium salts useful as fertilizers? From what sources are these salts obtained?

#### **Experiment 31. Preparation and Properties of Nitric Oxide.**

**Apparatus.** Four wide-mouthed bottles (6 to 8 oz.); two-holed rubber stopper to fit wide-mouthed bottle; thistle tube; delivery tube; sink or deep dish for water; oxygen generator, consisting of wide-mouthed bottle (6 oz.) with a two-holed rubber stopper carrying a delivery tube and a funnel

having a glass rod with one end ground into the funnel so as to form a stopper; test tube.

**Material.** Copper (wire, rivets, or turnings); concentrated nitric acid; sodium peroxide.

*a. Preparation.*

Arrange a wide-mouthed bottle with a stopper carrying a thistle tube and a delivery tube. Place in the bottle about 10 g. of copper and cover with a test-tubeful of water.

Pour about one third of a test-tubeful of concentrated nitric acid through the thistle tube of the generator and wait for the action to start. Collect the gas by displacement of water. To maintain the action in the generator, add from time to time small quantities of the concentrated nitric acid.

Note the color of the gas that first appears in the generator.

*Why does it not appear in the collecting bottle?*

*What is the difference in color between this gas and the one that does collect in the bottle?*

The gas that collects in the bottle over water is nitric oxide, NO.

Collect one full bottle, and another bottle half full of the nitric oxide, and let them stand in the dish of water for use later.

Note the color of the liquid in the generator. This color is characteristic of the water solution of cupric salts.

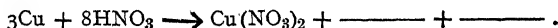
*What gas is often produced by the action of an acid with a metal?*

*What oxidizing action prevents our getting this gas here?*

*Name the product that is formed instead.*

The removal of oxygen from certain nitric acid molecules results in their reduction to nitric acid and water.

*Complete the equation :*



*b. Action with oxygen.*

Obtain from the instructor one of the bottles for generating oxygen by the action between water and sodium peroxide :



Gently loosen the stopper in the funnel so as to allow a *few drops* of the water in the funnel to fall upon the sodium peroxide at the bottom of the generator.

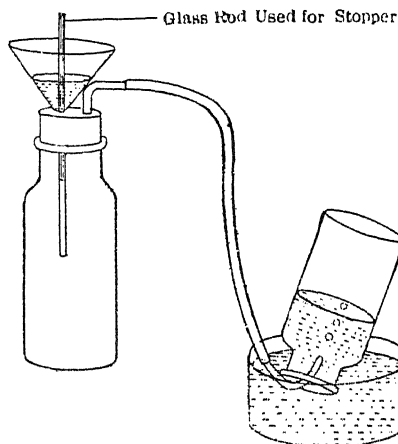


Fig. 49

Allow the air in the delivery tube to be displaced by the oxygen. If the action stops, let another drop of water fall on the peroxide.

*Caution! To avoid violent action, use water in small amounts with intervals between.*

Pass oxygen, a little at a time, by means of the delivery tube, into the bottle that is half full of nitric oxide, which is standing in the deep dish of water.

*What are the noticeable results?*

*In what respect does this colored gaseous product differ in composition from the nitric oxide?*

Complete the equation:



*c. Action with air.*

Allow the full bottle of nitric oxide to escape into the air.

*What evidence of a chemical change do you observe?*

*With what gas in the air does the nitric oxide combine?*

*Account for the reddish-brown gas that appeared in the generator at the beginning of part a.*

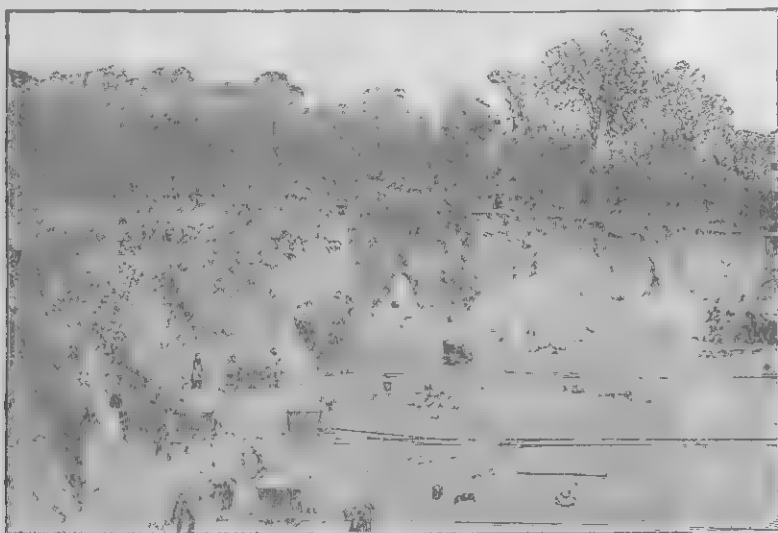
*Compare the solubility in water of nitric oxide and of hydrogen peroxide.*

*Make a drawing of your apparatus.*

## CHAPTER XXIV

### ELEMENTS OF THE NITROGEN GROUP

Nitrogen belongs to the same group or *family* with four other well-known elements, namely, phosphorus, arsenic, antimony, and bismuth. There is a general gradation in physical as well as chemical characteristics. As the atomic weights increase from nitrogen to bismuth, the elements change into elements entirely metallic in appearance. In this family of elements, nitrogen, the one having the smallest atomic weight,



From *National Geographic*

Fig. 50. THE PHOSPHATE QUARRY ON CHRISTMAS ISLAND, OF THE SOUTH SEAS

This flat summit of a submarine mountain, rising 15,000 feet from the floor of the Indian Ocean, was uninhabited when Europeans arrived, thirty years ago, and discovered its phosphate deposits.

is a colorless gas, that with the next higher is phosphorus, an almost colorless solid, easily fused and easily burned, the next, arsenic, is a grayish-white solid with the appearance of a metal and the chemical behavior of a nonmetal. Antimony and bismuth are metals both in their physical and chemical properties; while antimony shows a non-metallic nature in some of its compounds, bismuth is always a metal.

The melting points of the elements in this family increase with their atomic weights.

Phosphorus never occurs in nature in the free state. It is, however, quite widely distributed in small quantities. It occurs as calcium phosphate,  $\text{Ca}_3(\text{PO}_4)_2$ , in phosphate rock. Small amounts of phosphates also occur in all fertile soils, in iron ores, and in many granitic rocks, clays, etc. Bones consist of calcium phosphate to the extent of about 80 per cent. Phosphorus is an important and indispensable part of the tissues of all plants and animals. In the nerves, brain, blood, muscles, and in fact in all animal parts that are concerned in locomotion or reproduction, phosphorus is found. In seeds phosphorus is always present, particularly in the embryos, and without phosphorus plants cannot grow, hence its importance in the soil. As animals live these complex phosphorus compounds are gradually oxidized, phosphates being formed, which are then eliminated by the kidneys. Thus urine always contains phosphates, especially calcium and potassium phosphates. When these are then returned to the land as fertilizer, they are again taken up by the roots of plants and converted into complex phosphorus compounds in their tissues and seeds. These in turn are eaten by animals, and so the *phosphorus cycle* completes itself. The most important uses of phosphorus are in the manufacture of poison for rats and vermin, of matches, and of commercial fertilizer.

*Arsenic* is a steel-gray-to-black, brittle substance having a metallic luster. Arsenic compounds are used almost entirely for vermin and insect pests and are very poisonous. The two compounds of arsenic which are at present most commonly employed are *Paris green* and *lead arsenate*.

*Antimony* is a brittle metal. It is used in making *type metal* and *Babbitt metal* for machine bearings. As molten antimony solidifies, it expands and fills the molds, hence it is valuable in many alloys; moreover, it is hard and when mixed with lead and tin forms an alloy which is particularly suitable for type. *Bullets* and *shot* consist of lead alloyed with from 0.2 to 0.4 per cent arsenic.

*Bismuth*, too, is a brittle metal. It is used in preparing low-melting alloys which are employed in making readily fusible plugs in the pipes of automatic sprinklers for fire protection. So, for instance, Wood's

metal, consisting of 1 part tin, 2 parts lead, 1 part cadmium, and 4 parts bismuth melts at  $60.5^{\circ}\text{C.}$ , much below the temperature at which water boils. Bismuth compounds are used in medicine, the most common preparation being *bismuth subnitrate*,  $\text{BiO}\cdot\text{NO}_3\cdot\text{BiO}\cdot\text{OH}$ , a white powder which is difficultly soluble in water and practically tasteless. It is prescribed in cases of dysentery and other disturbances of the alimentary canal.

#### QUESTIONS FOR STUDY

1. Discuss the rôle of phosphorus in plant and animal life.
2. How much phosphoric acid,  $\text{H}_3\text{PO}_4$ , could be produced from one ton of calcium phosphate,  $\text{Ca}_3(\text{PO}_4)_2$ , the essential compound in phosphate rock?
3. What are the commercial compounds of arsenic used for?
4. What are antimony and bismuth, and what use is made of them?

#### Experiment 32. Cobalt Nitrate Tests.

**Apparatus.** Plaster block, made by pouring a thin mixture of plaster of Paris and water into a form on a glass or stone slab and cutting the mass just before it hardens into strips of suitable size; blowpipe; forceps; Bunsen burner.

**Materials.** Zinc sulphate; alum; magnesium sulphate; cobalt nitrate solution; unknowns.

a. Put a little of some zinc compound, as zinc sulphate in a cavity made in a plaster block with the top of a pair of forceps. Heat it as hot as possible at the end of a small blowpipe flame. Allow the residue to cool.

*Record the color of the residue in the table.*

Moisten the residue with a drop or two of cobalt nitrate solution. Again heat it intensely, and on cooling note the color of the mass that remains on the plaster block.

*Record in the table the color obtained.*

b. Place some alum in a new hole in the block. Repeat, as with the zinc sulphate, so as to obtain a characteristic coloration.

*Record results in the table.*

c. Make a new cavity in the block and place in it some magnesium sulphate. Heat the sulphate with the blowpipe until it glows brightly. Moisten the melted mass with a *very little* cobalt nitrate. Blow very



vigorously, and allow the mass to cool. Note carefully the delicate coloration.

*Record the colors obtained in the table.*

d. Obtain from the instructor an unknown compound, and test it with the blowpipe and cobalt nitrate, to determine whether it is a compound of aluminium, magnesium, or zinc.

*What color was obtained the second time heated?*

*What metal did the compound contain?*

TABLE

COMPOUND TAKEN	COLOR OF RESIDUE AFTER FIRST HEATING	COLOR OF RESIDUE AFTER HEATING WITH COBALT NITRATE

## CHAPTER XXV

### THE HALOGENS

The elements fluorine, chlorine, bromine, and iodine are called the *halogens*. They never occur in nature in the uncombined state, but are frequently met in compounds.

*Fluorine* is found chiefly as calcium fluoride, *fluor spar*,  $\text{CaF}_2$ , but it is also fairly widely distributed in minute amounts in granitic rocks, being a minor constituent of the mineral *apatite*, which consists essentially of calcium phosphate and a small percentage of calcium fluoride. Fluorine is present in extremely small quantities in soils, from which it gets into plants, and so into animals. In the enamel of the teeth of the latter notable amounts of fluorine are always present.

Hydrofluoric acid is poisonous. The gas and also the aqueous solution of the latter are used for *etching glass*, "diamond ink." Hydrofluoric acid is also used in the laboratory for the purpose of decomposing silicates in chemical analysis. Fluorine itself may be obtained by passing the electric current through a solution of potassium fluoride and hydrofluoric acid in water. Fluorine is a light-colored greenish yellow gas of extremely pungent odor. It unites with all elements directly, except with oxygen. It is probably the most active of all the chemical elements.

*Chlorine*, as we have seen and studied in an earlier chapter, is found mainly in common salt.

*Bromine* is found mainly as sodium bromide,  $\text{NaBr}$ , in connection with sodium chloride. It may be prepared by methods that are entirely similar to those used in making chlorine. Bromine forms bromides by direct union with many of the elements. Potassium bromide,  $\text{KBr}$ , is used in medicine and photography. It is used as an antiseptic, in making aniline dyes, bromides, etc. Bromine turns starch paste yellow.

*Iodine* is found in the ashes of seaweeds, also as sodium iodate,  $\text{NaIO}_3$ , in connection with Chile saltpeter. Iodine is a crystalline,

grayish-black solid having almost a metallic luster. It dissolves sparingly in water, but copiously in alcohol. The alcoholic solution is called tincture of iodine. It is used in medicine as an antiseptic. With the metals and some of the other elements iodine forms iodides. Thus we have sodium iodide ( $\text{NaI}$ ), potassium iodide ( $\text{KI}$ ), silver iodide ( $\text{AgI}$ ), calcium iodide ( $\text{CaI}_2$ ), phosphorus iodide ( $\text{PI}_3$ ). From sodium iodide, iodine may readily be obtained by treatment with sulphuric acid.

Iodine turns starch paste blue. The fact is used in testing for starch.

The thyroid gland contains iodine in the form of a complex compound, thyroiodine. In the treatment of goiter and other diseases connected with the thyroid gland, iodine compounds, like potassium iodide, or extract of the thyroid gland of the sheep, are frequently prescribed by the physician.

#### QUESTIONS FOR STUDY

1. Name the halogens, and one important compound of each found in nature.
2. Write the equations for the preparation of (a) hydrogen fluoride, (b) bromine, (c) iodine, (d) silver bromide. Write (d) in ionic form.
3. What use is made of (a) chlorine? (b) bromine? (c) iodine?

#### Experiment 33. Preparation and Properties of Bromine.

**Apparatus.** Ring stand; clamp; four test tubes; one-holed stopper and a delivery tube; beaker; Bunsen burner; test tube rack.

**Materials.** Potassium bromide; manganese dioxide; sulphuric acid, 2 to 1; carbon disulphide; chlorine water.

*Caution! Keep flames away from carbon disulphide. Its vapor is explosive when mixed with air.*

##### a. Preparation.

On a piece of paper, mix 1 gram of potassium bromide with an equal bulk of manganese dioxide.

Fit a test tube with a one-holed stopper carrying a delivery tube. Pour about 3 c. c. of sulphuric acid (2 to 1) into the test tube, and add

the mixture of potassium bromide and manganese dioxide. Clamp the tube containing the mixture so that the delivery tube shall extend to the bottom of an empty test tube, standing in a beaker of water. Warm the test tube containing the mixture *very gently*.

*What is the color of the bromine vapor and of the liquid bromine?*

Nearly fill the test tube containing the bromine with water. Save the mixture for future use.

*Is bromine heavier or lighter than water?*

*State reason for your answer.*

*What compounds would we expect to have formed by the action of potassium bromide with sulphuric acid?*

b. *Solubility of bromine.*

Add a few drops of carbon disulphide to a test tube one third full of water, and shake the contents.

*Are the two liquids completely soluble in each other?*

*How much does bromine dissolve in water?*

c. Add a few drops of the bromine water obtained in part *a* to the mixture of carbon disulphide and water.

Shake the resulting mixture vigorously and then allow it to stand for a short time.

*What color is given to the carbon disulphide by the bromine?*

*Does all the bromine dissolve in the carbon disulphide?*

*Is bromine more soluble in water or in carbon disulphide?*

d. *Test for a bromide.*

Dissolve a small crystal of potassium bromide in 2 or 3 c. c. of water, add a little carbon disulphide, and shake the mixture. Save the resulting mixture for part *e*.

*Does bromine, when combined with other elements, color carbon disulphide?*

*In what state must the bromine be to give the test with carbon disulphide?*

e. To the solution obtained in part *d* add a few drops of chlorine water and shake the mixture.

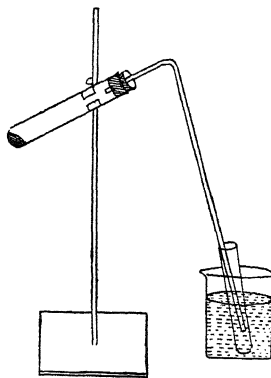
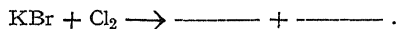


Fig. 51.

*What evidence is there that bromine has been set free by the action of the chlorine?*

*Complete the equation :*



*What name is given to this type of chemical reaction?*

*Give a test for a bromide.*

## CHAPTER XXVI

### CARBON AND ITS OXIDES

Like oxygen, the element carbon is found both free and combined. Free carbon is familiar as the black solid that makes up the greater part of the common substances, *coal*, *charcoal*, *coke*, and *lampblack*. The soft, slippery, shiny solid called *graphite* is also largely carbon. And, strange as it may seem, the valuable gem called *diamond* is also carbon — pure crystalline carbon.

Coal, charcoal, and coke are fuels. Lampblack is made into printing ink and black paint. The various varieties of these three forms

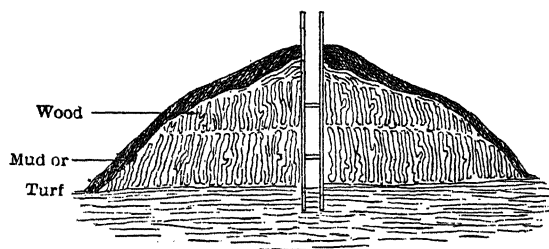


Fig. 52. Making Charcoal

A vertical section through the center of a pile.

of carbon are used in many chemical industries, e. g., the manufacture of iron and steel. Graphite is made into the "lead" of lead pencils, stove polish, lubricants, crucibles (for melting metals), and electrodes (for electric furnaces and electrolytic

cells). Impure diamond is used to polish pure diamond and other gems, to cut glass, and as the cutting part in the diamond drill that is used to bore wells for oil and water and to drill hard rocks.

The natural and manufactured compounds of carbon number over two hundred thousand. These compounds are so numerous that they are treated by themselves in a branch of chemistry called "organic chemistry." Carbon is the fundamental element of plant and animal life. Compounds composed of carbon, hydrogen, and oxygen and in some cases also of nitrogen, form such common substances as sugar, starch, fat, cotton, paper, flour, rubber, soap, wool, and meat. With hydrogen it forms a large class of compounds called hydrocarbons which are found in illuminating and fuel gases, petroleum, kerosene

gasoline, lubricating oils, wax, paraffin, and turpentine. The manufactured compounds of carbon include dyes, medicines, perfumes, and a vast number of other substances.

The commonest inorganic compounds—those usually considered in the branch of chemistry we are studying, are the carbonates and the oxides. The carbonates are compounds of carbon, oxygen, and a metal such as calcium, magnesium, or sodium. Thus, calcium carbonate ( $\text{CaCO}_3$ ), is the natural substance called limestone, marble, or chalk. Sodium carbonate ( $\text{Na}_2\text{CO}_3$ ) is the common substance, washing soda. There are two carbon oxides—carbon dioxide ( $\text{CO}_2$ ) and carbon monoxide ( $\text{CO}$ ).

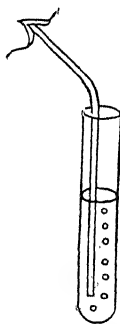


Fig. 53. Blowing Through Limewater to Show the Presence of Carbon Dioxide in the Breath

We have already seen that carbon dioxide is formed when carbon burns in oxygen and also when a candle burns in air. These are examples of a common chemical change. The process of burning or combustion consists usually in the union of carbon (and also hydrogen, if present) with oxygen. This means that carbon dioxide is being constantly formed by the burning of such common fuels as wood, paper, coke, coal, charcoal, oil, and gas. In fact, carbon dioxide is always one of the products of combustion yielded by burning any substance which contains carbon, e. g., sugar, starch, wax, meat, milk, camphor, alcohol, oil, dyes, fat, and drugs. The presence of carbon dioxide can be shown by bubbling the products of combustion, e. g., smoke, through limewater.

One of the products of the digestion of food is carbon dioxide. This means that the element carbon, which is one of the constituents of our food, is oxidized to carbon dioxide. Heat is liberated when carbon is oxidized and this chemical change is one part of the continuous and complicated processes by which our bodies keep uniformly warm.

Carbon dioxide is also formed by other chemical changes, such as the decay of many kinds of animal and vegetable matter and fermentation

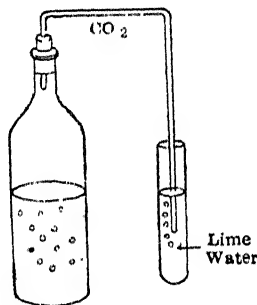


Fig. 54. Carbon dioxide evolved from soda water causes a precipitate in lime-water.

of organic substances like sugar. The latter process is illustrated by the liberation of carbon dioxide during the raising of bread.

Carbon dioxide does not assist combustion nor does it burn for it is oxidized as much as it can ever be oxidized. It is not poisonous, though the presence of a small quantity in the air of a room is objectionable. The carbon dioxide that is exhaled from our lungs is one of the products formed by the oxidation of the tissues of the body, new tissue itself being formed from the food. The carbon needed for the rebuilding of tissue is supplied by starch and other substances we eat. Carbon dioxide is a waste product of animal life. Plants absorb carbon dioxide and transform it into starch, whereas animals eat starch as food, assimilate it, and oxidize the carbon to carbon dioxide, which is exhaled into the atmosphere ready for the plants again.

Carbon monoxide is a compound of carbon and oxygen, but the two compounds differ in properties and composition. They can be prepared from each other by chemical processes.

Carbon monoxide, like the dioxide, is a gas without color, odor, or taste, but in other properties the two gases differ. Thus carbon monoxide is only very slightly soluble in water and does not form a compound with water. It is lighter than carbon dioxide; a liter weighs 1.25 gms. It is a dangerous poison because the lack of odor prevents its detection. Many deaths have been caused by breathing air containing it. A small fraction of a per cent in the air causes a stupefying gas. Gases from charcoal stoves used for heating purpose in the winter contain carbon monoxide and many deaths have occurred from their use in small and ill-ventilated rooms.

Not only does carbon monoxide unite readily with oxygen, but it withdraws oxygen from hot oxides; carbon itself acts in the same way. In chemical language, carbon monoxide is a *reducing agent*, i. e., it withdraws oxygen from compounds. In the manufacture of iron from iron ores, the ore, which is an oxide ( $\text{Fe}_2\text{O}_3$ ) is reduced by carbon monoxide in a blast furnace, the gas for this purpose being produced by the incomplete combustion of coke.



## QUESTIONS FOR STUDY

1. In what form does free carbon occur in nature? Name ten familiar solids, three liquids, and two gases which contain carbon.
2. How is carbon dioxide related to plant and animal life?
3. What would we mean by the "carbon cycle"?
4. Suggest a method for proving that all the various forms of carbon described are really carbon.

**Experiment 34. Preparation and Properties of Carbon Dioxide.**

**Apparatus.** Flask with two-holed stoppers carrying thistle tube and delivery tube; three wide-mouthed bottles; three glass plates; three test tubes; glass tube; sink of water.

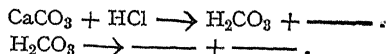
**Material.** Marble chips; concentrated hydrochloric acid; limewater; blue litmus solution; candle; wooden splinter.

*a. Preparation.*

Into the flask fitted with delivery tube and thistle tube reaching nearly to the bottom of the bottle, put marble chips to about the depth of one inch. Cover the marble with water, and add concentrated hydrochloric acid, a few c. c. at a time, so as to get a *moderate* action. Collect three bottles of gas by the displacement of water and save for use in parts *a*, *b*, *d*, and *e*.

*Describe the action in the generator.*

*Complete the equation:*



*Is carbonic acid stable at ordinary temperatures?*

*Why can any of the common acids be used in preparing carbon dioxide?*

*b. Odor and color.*

Inhale some of the gas from the bottle.

*What is the effect of the gas on the nose?*

*Has carbon dioxide color?*

*c. Solubility.*

Let the gas from the generator bubble through a test tube half full of blue litmus solution.

*What change takes place?*

*Why is carbon dioxide called an acid anhydride?*

*How soluble is carbon dioxide in water?*

*How can the solubility of the gas be increased in the preparation of effervescent drinks?*

*d. Relative weight and relation to combustion.*

Slowly invert a bottle of carbon dioxide over a lighted candle.

*Results?*

*What three properties of carbon dioxide are shown?*

*e. Action with limewater.*

Pour half a test-tubeful of limewater into a bottle of carbon dioxide and shake the bottle

*Result?*

*Write the equation to show the formation of the precipitate, calcium carbonate.*

*f. Product of combustion.*

Burn a wooden splinter in a bottle of air. Add limewater and shake the bottle.

*Result?*

*How does this show that the splinter contains carbon?*

Blow through a glass tube into a test tube half full of limewater.

*Result?*

*What gas do we breathe out from our lungs?*

*How is it produced in the body?*

**Experiment 35. Preparation and Properties of Carbon Monoxide.**

**Apparatus.** Ring stand with 1 ring and 1 clamp; wire gauze with asbestos center; beaker, 200 c.c.; side-arm test tube, with single-holed rubber stopper to fit; small thistle tube; delivery tube; enameled pan; 2 wide-mouthed bottles, 4 oz.; glass plate; Bunsen burner.

**Material.** Formic acid, sp. gr. 1.2; concentrated sulphuric acid; limewater.

*a. Preparation.*

Pour 5 c. c. of formic acid and an equal amount of concentrated sulphuric acid through the thistle tube into the side-arm test tube.

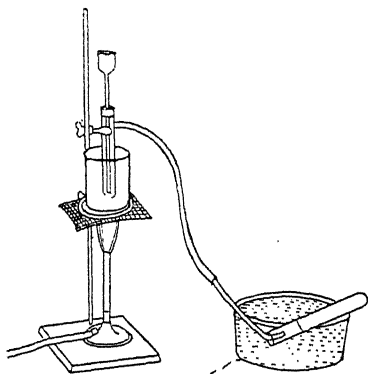


Fig. 55

Heat the water in the beaker till it boils. Collect the gas in small bottles as long as it is evolved. *Do not let carbon monoxide escape into the room. It is very poisonous.* Get rid of the gas in any of the bottles that you do not need by burning it.

*b. Properties.*

Burn a bottle of carbon monoxide.

*Describe the color of the flame.*

*Compare this flame with that of burning hydrogen.*

Pour 15 c. c. of limewater into the bottle in which the carbon monoxide has been burned and shake the liquid around gently.

*Result?*

*What gas is formed by the burning of carbon monoxide?*

*Write the equation for the reaction.*

*c. General questions.*

The formula of formic acid is  $\text{HCHO}_2$ .

*Write the equation for its decomposition.*

*Explain how concentrated sulphuric acid helps to bring about this decomposition.*

## CHAPTER XXVII

### GASEOUS AND LIQUID FUELS

Gaseous fuels include natural gas and the various mixtures obtained from coal, e. g., producer gas, water gas, and coal gas. *Natural gas* exists in the earth and is usually found in regions where petroleum is found. It is obtained by boring wells. It is about 90 per cent methane ( $\text{CH}_4$ ), which is the chief heat-producing constituent. It is used as fuel to heat houses and to generate steam in many industries, e. g., making steel, glass, brick, and pottery, in places where it is available. As a fuel it is cheap and efficient.

*Producer gas* is made by forcing air (sometimes together with a little steam) through a deep coal fire in a special kind of furnace. The gaseous product contains 25 per cent to 30 per cent of carbon monoxide, 50 per cent to 60 per cent of nitrogen, and 10 per cent to 13 per cent of hydrogen (if steam is used). The heat value is low on account of the large proportion of nitrogen, but the cost is not great because the gas can be quickly made and efficiently burned. Producer gas is used as a fuel in gas engines and in many industrial and metallurgical operations.

*Water gas* is made by forcing steam through a mass of hot coke or anthracite coal. It is essentially a mixture of hydrogen and carbon monoxide, and is used to some extent as fuel. More often it is enriched by spraying in petroleum oil and then used as an illuminant.

*Coal gas* is made by heating coal in closed retorts and purifying the volatile product. It is largely hydrogen (45 per cent to 50 per cent) and methane (30 per cent to 36 per cent); other ingredients are hydrocarbons, carbon monoxide and dioxide, and nitrogen.

*Acetylene* ( $\text{C}_2\text{H}_2$ ) is a gas which is prepared by the interaction of calcium carbide and water. Acetylene dissolves in acetone, and cylinders in which considerable gas has been dissolved are used to furnish gas for mining camps and other localities where ordinary illuminating gas is not available.

A mixture of acetylene and considerable oxygen burns with an intensely hot flame, a temperature of nearly  $3,000^{\circ}\text{C}$ . can be reached if the right mixture of acetylene and oxygen is burned in a special blow-pipe called an oxyacetylene torch. Ordinary tools cut hard metal slowly, but the tip of the oxyacetylene flame when passed slowly across the metal melts ("cuts") it quickly. Metal structures, such as fences, bridges, frames of buildings, "scrapped" warships, etc., are speedily dismantled by this flame. The oxyacetylene flame is also extensively used for welding metals.

Three products from *petroleum* — fuel oil, gasoline, and kerosene—are used as fuels. They are mixtures of hydrocarbons. In burning, these compounds are decomposed; the carbon forms carbon dioxide and the hydrogen forms water.

*Fuel oil* is extensively used on warships and steamships and in many manufactories. By using oil instead of coal to generate steam, additional space is provided for cargo, more efficient combustion is attained, speed may be greatly increased, and the cruising radius enlarged, to say nothing of the labor saved and the cleanliness secured.

*Gasoline* is the fuel used in the engines of automobiles, trucks, motor boats, motor cycles, and airplanes. Gasoline is very volatile and the vapor burns readily. If the vapor is mixed with air and the mixture is ignited by an electric spark, the combustion is so rapid that it is practically an explosion; the gases, suddenly expanded, exert pressure, which is converted by the machinery into steady and continuous motion.

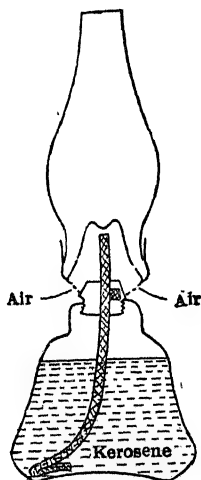


Fig. 56. Diagram of a Kerosene Lamp

*Kerosene* is used as a fuel to a limited extent in engines and cooking stoves. In some portable stoves kerosene is burned by means of a large wick—as in the kerosene lamp, which has displaced the old-fashioned Chinese vegetable-oil lamp in many localities.

Different kinds of *alcohol* are used as fuel, usually on a small scale. They burn without smoke and have a high heat value. *Methyl alcohol* ( $\text{CH}_3\text{OH}$ ) is also called wood alcohol. It is very poisonous when taken internally.

*Ethyl alcohol* ( $C_2H_5OH$ ), also called grain alcohol, is made from the starch obtained from corn and potatoes.

## QUESTIONS FOR STUDY

1. Suggest a way in which natural gas may have been formed.
2. Why does the use of the bellows on the blacksmith's forge cause a more intense heat?
3. Name 5 fuels and 3 illuminants.
4. Explain the use of gasoline in an automobile.

**Experiment 36. Destructive Distillation.**

**Apparatus.** Hard glass test tube; two sets of stoppers and delivery tubes; two 8-in. test tubes; ring stand; clamp; Bunsen burner.

**Material.** Wood (splinters); soft coal; litmus paper.

*a. Destructive distillation of wood.*

Arrange apparatus having the lower end of the delivery tube half of an inch from the bottom of the condensing tube. Fill a hard glass test tube with splinters of wood. Heat, gently at first, and then strongly, until no further change can be noted.

*Describe the appearance of the volatile matter passing off from the wood.*

While heating, bring a flame to the end of the jet tube.

*Result?*

When the action is complete, allow the apparatus to cool, then examine the contents of the test tube.

*What is the substance found there?*

*How does it differ from the original wood?*

*What is the appearance of the substance in the condensing tube?*

*Describe any distinct layers that can be distinguished.*

Test the liquid in the condensing tube with litmus paper.

*Result?*

*Note the odor of the liquid.*

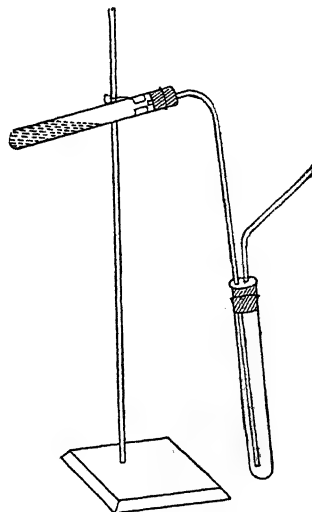


Fig. 57

This tarry distillate is a mixture called pyroligneous acid. It contains wood alcohol, acetic acid, acetone, tar, and other substances.

The decomposition by means of heat of a complex substance such as wood, into simpler substances, some of which are condensed to liquids, is called *destructive distillation*.

*Name three direct products of the destructive distillation of wood.*

*b. Destructive distillation of coal.*

Free the hard glass tube from charcoal, and half fill it with finely crushed soft coal. Replace the condenser and tubes with a fresh set, and heat as before.

*Describe the appearance of the volatile matter in this case.*

*Is it the same as that obtained from the wood?*

While the heating is going on, lay a piece of red litmus paper over the end of the jet tube.

*Result?*

*What kind of compound is shown to be present by this test?*

Bring a flame to the end of the jet tube.

*Result?*

Heat the tube until gas is no longer given off. Then allow the tube to cool. When you can handle the hard glass test tube, pour its solid contents on the base of the ring stand.

*Describe the residue as to color, structure, and weight as compared with the original coal.*

This residue is coke.

Examine the condensing tube.

*Describe the liquid deposited there.*

*Why is the distillation of coal a destructive distillation?*

*Name three direct products of this destructive distillation.*

*Make a drawing of your apparatus.*

## CHAPTER XXVIII

### COLLOIDS

If we take some very finely powdered substance, such as flour, and stir it up with water, we obtain an apparently uniform mixture. On standing, the solid will slowly settle, and the mixture will thus gradually separate into solid and liquid. Such mixtures are called *suspensions*. Muddy water is a common example of a suspension.

Many liquids will dissolve in one another in all proportions; for example, alcohol and water. These substances are said to be *miscible* in all proportions. If we mix water and carbon disulphide ( $\text{CS}_2$ ), we very soon notice a sharp line of separation between these two *immiscible* liquids. Some liquids, such as phenol ( $\text{C}_6\text{H}_5\text{OH}$ ) and water, have a limited solubility in each other, the solubility depending on the temperature.

If oil and water, two immiscible liquids, are violently shaken up together, small globules of the one may remain suspended in the other for a long time.

This is a mechanical mixture like a suspension, and is called an *emulsion*. Salad dressing is such an emulsion; we

make it by shaking olive oil and vinegar together. After long standing emulsions will separate into the two liquids.

The difference between solutions on one hand and emulsions and suspensions on the other is probably due to the size of the particles. In a solution we have molecules distributed throughout the liquid which are very small and move about rapidly. In suspensions and emulsions the particles are many thousand times larger than molecules and can be seen with a microscope.

Between the true solutions and suspensions and merging into both we have what is known as colloidal solutions, or *colloids*. Colloidal

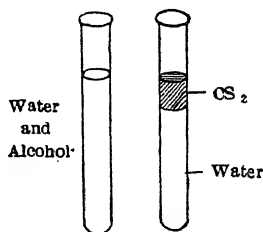


Fig. 58. Water and alcohol are miscible in all proportions; water and carbon disulphide are immiscible.



solutions usually look perfectly clear to the naked eye, but when a beam of light is thrown through them they can be seen to be not entirely homogeneous.

If we put some starch into cold water we find that it is not soluble but forms a suspension. But if we boil it in water the grains of starch swell and break, and the starch becomes finely diffused throughout the water. This clear liquid is a colloidal solution. When the liquid is poured through a filter paper, almost all the *starch goes through the paper*. On cooling, the liquid becomes a clear stiff jelly.

If we pass a beam of light through a solution of common salt, we see no distinct path because the molecules are too small to reflect light. But if we pass a beam of light through the clear colloidal solution of starch we see the bright path.

The size of the particles of the dissolved substance in a colloidal solution lies between the size of those in a true solution and those in a suspension. They are so small that they can move with some freedom, and they never settle out as do suspensions. On the other hand, they probably do not move around as freely or as rapidly as dissolved molecules as they are so much larger. Colloidal solutions exert a very small osmotic pressure and have almost the same freezing and boiling points as the pure solvents. This shows that the dissolved particles are very large.

The importance of the study of colloids in chemistry can be estimated when we realize that a very large part of living matter is made up of colloidal substances. The cells of plants and of animals; part of the blood; the sap of trees like rubber; much of our food for instance—these are typical examples of colloids. They are furthermore of great importance in the industries; the process of tanning leather and of dyeing, the making of artificial silk and of glass involve a knowledge of colloids.

Glue and gelatin well illustrate some of the properties of colloids. Colloidal solutions of a number of such simple substances can be prepared. Some of these, like glue, are very stable and can be preserved for a long time; others are unstable and on being treated with small amounts of certain substances or on being heated precipitate the colloidal material. This process is often spoken of as *coagulation*.

For example, the colloidal material in milk—casein—can be rapidly coagulated at 50° C. by the addition of a little acetic acid. This same change takes place slowly in the preparation of junket and in making cheese.

Some colloidal solutions on cooling or on evaporation turn into jellylike materials which are called *gels*. The most familiar example is gelatin, a 5 per cent solution of which turns into a gel at 18° C. Gels go back into colloidal solutions if we warm them or add more water. They all contain a great deal of the solvent although they are very rigid solids. For this reason a relatively small weight of a dry substance like gelatin will in solution produce a large amount of gel. All the fruit jellies which are made directly from the fruit are gels; a colloidal substance, pectin, which is present in the fruit, causes the gel to form. *Solid alcohol* is a gel formed by dissolving a colloidal substance (stearic acid) in denatured alcohol.

#### QUESTIONS FOR STUDY

1. What happens when milk stands for some time? Why doesn't this happen to a gelatin solution?
2. What common household substances are colloidal gels? How can they be liquefied?
3. The white of a raw egg is a colloidal solution. How could you coagulate it?

#### Experiment 37. Properties of Carbon.

**Apparatus.** Hard glass test tube with one-holed stopper and bent delivery tube; two test tubes; beaker, 200 c. c.; evaporating dish; funnel; ring stand with clamp, and one small ring to support the funnel; stirring rod; Bunsen burner; small sheet of paper.

**Material.** Copper oxide, powdered; charcoal, powdered; limewater; bone-black; sugar; copper sulphate solution, 1 to 40; filter paper; cider vinegar.

##### *a. Carbon as a reducing agent.*

Arrange the apparatus—with the hard glass test tube (fitted with stopper and delivery tube) *almost* horizontally clamped to the ring stand, and the delivery tube extending into the test tube containing the limewater.

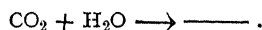
Fill one tenth of a test tube with copper oxide, then pour the oxide on a sheet of paper. Using the same test tube, measure an equal volume of powdered charcoal. Add the charcoal to the copper oxide on the piece of paper and mix the two thoroughly. Pour the mixture into the hard glass test tube. Pour limewater into the test tube into which the delivery tube extends until the limewater just touches the end of the delivery tube.

Heat the hard glass test tube, cautiously at first, commencing at the part around the upper portion of the mixture, and gradually moving the flame toward the closed end of the tube.

*What change do you observe in the limewater at first?*

Carbon dioxide is the anhydride of carbonic acid.

*Complete the equation:*



*What base is in solution in the limewater?*

*Write the equation for the neutralization of this base with carbonic acid.*

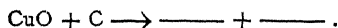
Allow the tube to cool, and when cold pour its contents into a 200 c. c. beaker. Let a small stream of water flow into the beaker.

*What substance is carried away by the water?*

*What is left in the beaker?*

*What element is taken from the copper oxide?*

*Complete the equation:*



*What name is given to the process of removing oxygen from a compound?*

*What kind of agent is carbon in this case?*

b. *Carbon as a decolorizer.*

Thoroughly mix a little boneblack with some vinegar in an evaporating dish. Fold a sheet of filter paper and place it in the funnel. Pour boneblack on the filter and scoop out a hollow in the center of the boneblack. Into this hollow, pour the mixture of vinegar and boneblack. Collect the filtrate and note its color.

*Result?*

Put enough sugar into a test tube to fill the rounded part. Heat this slowly and evenly, rotating the tube as the sugar melts. The sugar should be well browned, but not burned. The product contains caramel, a substance used as a flavor in cooking. As soon as the tube containing the caramel has cooled, fill two thirds of it with water, and warm the mixture until the solid has dissolved. The solution thus obtained is used to illustrate the impure sugar solution of a sugar refinery. The heat used in obtaining crude sugar helps to give the product a brown color which must be removed to get white sugar.

Filter the solution of sugar and caramel as you did the vinegar.

*Result?*

Filter a dilute solution of copper sulphate through boneblack.

*Result?*

*Can the color be removed from all liquids by filtering them through boneblack?*

Make a drawing of your apparatus.

## CHAPTER XXIX

### SILICON AND BORON

The element silicon does not occur free in nature but its compounds are very abundant and widespread, especially silica and the silicates. Sand is chiefly silica, while clay and most rocks consist of one or more silicates. Silicon (through its compounds) makes up about one fourth of the earth's crust, being next to oxygen in abundance.

The commonest compound of silicon is the oxide ( $\text{SiO}_2$ ) called silicon dioxide, or more usually *silica*; sand, sandstone, and certain other substances found in enormous quantities consist of silica. Its purest form is quartz, or rock crystal, which occurs in fine, clear, six-sided prisms. Some of the semiprecious stones, as amethysts, are impure forms of quartz.

Silica is also found in plants, especially in the leaves and stems, where it appears to give rigidity. Certain one-celled marine animals have skeletons containing silica. These animals belong to the *infusoria*; in some localities there are immense deposits of very fine sand called *infusorial earth*. This material is used to absorb nitroglycerin in making dynamite; also for removing coloring matters from oils and in making polishing powders.

Silica is a hard brittle substance and hence is used as an abrasive. *Sandpaper* consists of sharp quartz sand glued on paper. Silica can be melted in the oxyhydrogen flame and made into fine threads, tubes, flasks, or crucibles.

Apparatus made of fused quartz does not expand or contract to any appreciable extent even with considerable change in temperature. Quartz vessels are not attacked by the common acids. However, they are dissolved rapidly by hydrofluoric acid or concentrated alkali solutions.

*Water glass*, or sodium silicate ( $\text{Na}_2\text{SiO}_3$ ), is made by fusing together silica and sodium carbonate. A solution of sodium silicate is an excellent preservative for eggs since it fills the pores of the shells and

thus makes them air-tight. It is used to render wood or cloth fireproof, and as a cement for glass and pottery.

Silicic acid is precipitated when sodium silicate is acidified. The acid comes down as a gelatinous precipitate containing a rather indefinite amount of water. When heated sufficiently, the jelly loses all its water and forms silica. Silicic acid is a very good example of a colloid.

Common glass is made by melting together sand, sodium carbonate, and calcium carbonate. At a high temperature the carbonates react with the sand, forming glass, a mixture of sodium and calcium silicates. Such a mixture of silicates solidifies on cooling but does not crystallize. It stays as a transparent amorphous mass which may be considered as a supercooled solution of the one silicate in the other. On long standing, certain grades of glass begin to crystallize, and this results in an opaqueness and a fragility of the glass. Hard glass, which melts at a higher temperature than ordinary, or soft glass, consists of a mixture of potassium and calcium silicates. Flint glass contains the silicates of lead, barium, and potassium.

If the reaction between sand and carbon is carried out under certain conditions with a large excess of carbon, a compound of these two elements is formed, known as *carbon silicide*, or carborundum. The core of the furnace which is used is a mass of granulated coke. A mixture of sand, sawdust, coke, and a little salt is piled around the central core. The heat generated by the passage of a strong current through the coke is sufficient to cause the reaction to take place in the surrounding mass. Carborundum is a crystalline solid with a beautiful purple color. It is extremely hard and is used as an abrasive in the form of grinding wheels and hones, and as a fine polishing powder.

*Boron* occurs in nature in boric acid and its sodium and calcium salts. It is never found in the uncombined state. *Boric acid* has the composition  $\text{H}_3\text{BO}_3$ . It may be prepared from *borax*,  $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$ , by treating a hot, concentrated solution of the latter with either hydrochloric or sulphuric acid.

Boric acid and borax are the most important compounds of boron. Boric acid is also known as *boracic acid*. It dissolves in water. At room temperature, the saturated solution contains about 4 per cent. It is frequently used as an antiseptic wash, especially in treating the

eyes, for it is a very mild agent. Compresses of saturated boric acid solution are frequently applied on parts of the body that are swollen because of blood poisoning. It was formerly used to preserve milk, fish, etc., but not any more, as it is injurious to the health.

Borax is alkaline towards litmus, for it is a salt of a very weak acid and a powerful base. In the laundry borax is used for softening water. It is also used as an antiseptic, as a mordant in dyeing fabrics, and as a flux in welding iron and brazing metals together. In making glazes and enamels for pottery and enameled ironware, borax and boric acid are often used, also in the production of certain kinds of hard glass. The borax bead test is used in testing small quantities of minerals.

#### QUESTIONS FOR STUDY

1. Name five important uses for silica.
2. What is the difference between water glass and ordinary glass?
3. How does silica occur in nature?
4. What is borax? What use is made of it?

#### Experiment 38. Borax Bead Test.

**Apparatus.** Mounted platinum wire, or glass rod, five inches long; Bunsen burner; triangular file.

**Material.** Powdered borax; cobalt nitrate; manganese dioxide; chrome alum, or chromium sulphate; ferric chloride or other iron compound.

a. Bend the end of a platinum wire into the shape of a letter J, 2 mm. across the opening. Heat the loop red hot, and dip it into powdered borax. Heat again in the hottest part of the flame. The borax swells, loses its water of crystallization, and then melts to a transparent glass.

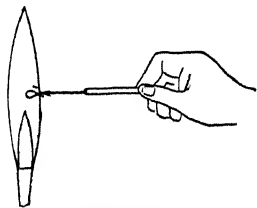


Fig. 59

In case a platinum wire is not available, heat the end of a glass rod in the flame. Then dip the hot end into powdered borax. Heat in the hottest part of the flame the borax that sticks to the rod, until the borax swells up to an irregular-shaped mass on the end of the rod.

*b.* Touch the hot irregular mass to a tiny bit of cobalt compound. Heat in the hot, outer portion of the Bunsen flame (oxidizing flame), until a clear glassy bead is obtained. Note the characteristic color and record in table.

To remove the bead from the platinum wire, heat it red-hot and quickly shake off the molten bead into the sink or waste jar. Make a fresh bead and examine it. If it is not colorless, repeat the operation.

To remove the bead from the glass rod, scratch it with a sharp triangular file, and break it off as you would cut a piece of glass tubing; or let a drop of water fall on the hot rod and tap the end of the rod on the table.

*c.* Touch a newly made bead while hot to a bit of manganese dioxide. Fuse as before. Note the characteristic color given by manganese compounds, and record in the table.

*d.* Repeat the same with a chromium compound. Record your result in the table.

*e.* Repeat with an iron compound in the oxidizing flame. Record your result in the table.

TABLE

COMPOUND TAKEN (NAME)	FORMULA	COLOR OF BEAD



## CHAPTER XXX

### CALCIUM AND ITS COMPOUNDS

Metallic calcium is made by the electrolysis of the fused salt. It has few commercial uses.

*Calcium carbonate* ( $\text{CaCO}_3$ ) is found in large quantities in nature as limestone, marble, and chalk. Iceland spar, or calcite, is a very pure crystalline form of calcium carbonate. *Calcium sulphate* ( $\text{CaSO}_4$ ) occurs as anhydrite, *gypsum* ( $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ ), and alabaster. In natural waters calcium bicarbonate and calcium sulphate are quite commonly found. They produce the hard sediment in teakettles, boilers, etc. These sediments generally also contain silica, oxides of iron, aluminium, magnesium, etc. These have been dissolved from the rocks or soils with which the waters have come in contact.

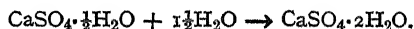
*Marble* and *limestone* serve as building stones, also for making lime, glass, and cement, and reducing iron ores. In the form of *chalk*, calcium carbonate is used as *whiting*, which, when ground up with linseed oil, forms *putty*.

*Lime*,  $\text{CaO}$ , is made by heating limestone or marble in kilns.



*Calcium chloride* is very soluble in water, and its brines are commonly used as the cold liquid that circulates in the pipes in refrigerators. This brine is kept cold by means of evaporation of ammonia which has been liquefied under pressure.

On carefully heating gypsum to about  $110^\circ$  it loses water and becomes  $\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$  which is known as *plaster of Paris*. When it is afterwards mixed with water, the whole hardens to a solid mass, which depends upon the fact that plaster of Paris unites with water and forms gypsum again.



Plaster of Paris is much used in making hard wall plaster, gypsum, calsomine, casts, and bandages in surgery.

As all plants and animals contain sulphur and calcium salts, calcium sulphate in the form of gypsum is commonly used as a fertilizer under the name of *land plaster*.

*Calcium nitrate* ( $\text{Ca}(\text{NO}_3)_2$ ) is very soluble in water. It is formed in composted manure and also on the whitewashed walls of stables. The ammonia present in the stable air is gradually oxidized to nitric acid, which then attacks the calcium carbonate on the walls, forming calcium nitrate and carbon dioxide. This, of course, destroys the wall coating. The calcium nitrate in the soil is a good plant food, for it contains nitrogen and calcium both in the available form.

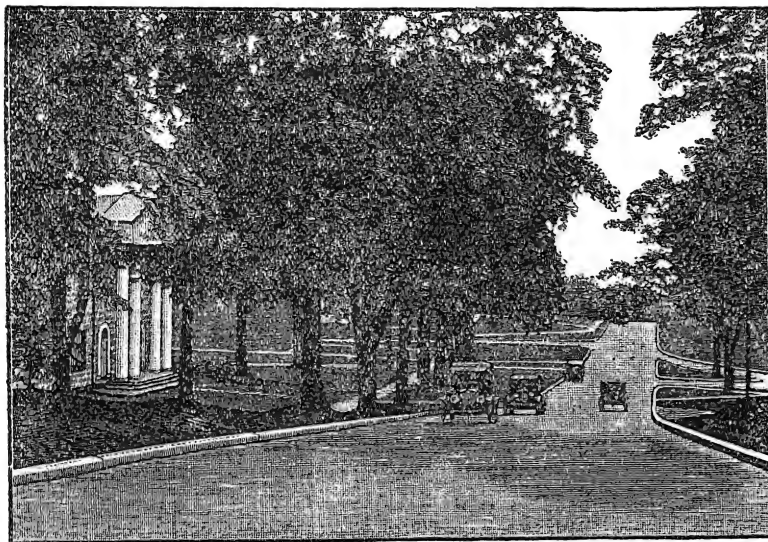


Fig. 60. The Beautiful Portland Cement Concrete Drive Through the Campus of Williams College, Williamstown, Massachusetts

*Calcium phosphate* ( $\text{Ca}_3(\text{PO}_4)_2$ ) is insoluble in water but soluble in dilute acids. It occurs in crystals as apatite, is the chief constituent of bones, and is found in phosphate rock, guano, and barnyard manure. It is a very important fertilizer.

*Portland cement* is made by heating limestone together with aluminium silicates in a kiln, and then grinding up the clinkers to a fine powder. When mixed with water it will unite with the latter, forming a hard stonelike mass. Portland cement sets, even under water, and



**Experiment 39. Hard Waters.**

**Apparatus.** Carbon dioxide generator, consisting of wide-mouthed bottle with two-holed stopper to fit, carrying thistle tube and delivery tube for leading gas to bottom of a test tube; four test tubes; small funnel; stirring rod; Bunsen burner.

**Material.** Calcium sulphate (plaster of Paris); magnesium sulphate; marble chips; dilute hydrochloric acid, 1 to 4; limewater; distilled water; filtered soap solution; filter paper.

a. Add a drop or two of soap solution to distilled water in a test tube and shake the tube.

*Are lasting suds produced?*

This result is characteristic of "soft" waters.

Drop a pinch of magnesium sulphate into a test tube two thirds full of water. Close the mouth of the tube with the thumb and shake it thoroughly. Add a few drops of the soap solution. Shake the tube.

*Do suds form?*

*What is produced?*

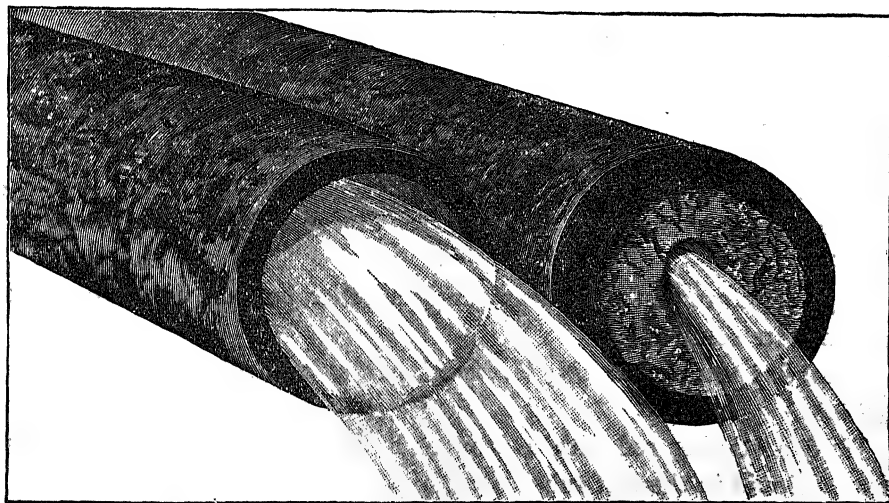


Fig. 62. a. Soft Water Flowing Through a Pipe. b. Hard Water Flowing Through a Pipe Depositing Insoluble Substances on the Sides

This is an insoluble magnesium soap. Waters that behave in this way with soap solution are "hard" waters.

Continue the addition of soap solution until suds finally form.

*Why does it cost more to wash with hard water than with soft water?*

*What would tend to become entangled in the fibers of a fabric when clothes are washed with hard water?*

b. Put some pieces of marble into a generator and cover them with dilute hydrochloric acid. Lead the carbon dioxide formed to the bottom of a test tube two thirds full of limewater.

*What is the first result?*

The insoluble compound formed is calcium carbonate.

*Write the equation for its precipitation.*

*What happens to the precipitate on the continued passing of the carbon dioxide?*

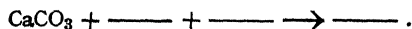
The compound formed is calcium bicarbonate,  $\text{CaH}_2(\text{CO}_3)_2$ .

*Is it soluble or insoluble?*

*What acid was formed when the excess of carbon dioxide reacted with the water in the test tube?*

*What effect did this acid have on the precipitate, calcium carbonate?*

*Complete the equation:*



c. Pour a little of the water solution of calcium bicarbonate into another test tube. To this portion, add a few drops of soap solution and shake the tube.

*What kind of water is the water containing calcium bicarbonate in solution?*

In a separate test tube heat gently another portion of the calcium bicarbonate solution.

*What happens on warming the liquid, particularly near the walls of the test tube?*

Hold a stirring rod with a drop of limewater on its end above the heated liquid in the test tube.

*What is the effect on the limewater?*

*What does this show?*

Continue heating the liquid in the test tube till it boils.

*What is the second effect of heat on a water solution of calcium bicarbonate?*

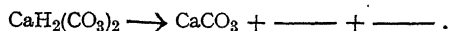
Filter the contents of the tube just heated. Test the filtrate with a few drops of soap solution.

*Has the water solution of calcium bicarbonate been softened?*

*Explain.*

Such a hard water is called a "temporary" hard water.

*Complete the equation for the softening of a temporary hard water by boiling:*



*Account for the crust formed on the inside of teakettles in which temporary hard water is boiled day after day.*

d. Add a pinch of calcium sulphate (plaster of Paris) to a test tube two thirds full of water. Shake the tube, filter the contents, and divide the filtrate into two portions. Using one, determine with soap solution whether or not the water is hard.

*Result?*

Boil the second portion, and then test with soap solution.

*Result?*

*Is the water solution of calcium sulphate softened by boiling?*

Such water is "permanent" hard water.

*Would permanent hard water form a deposit on a teakettle?*

*Explain.*

A water often contains both temporary and permanent hardness.

*How could you show the presence of each kind of hardness in the presence of the other?*

## CHAPTER XXXI

### MAGNESIUM, ZINC, AND MERCURY

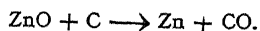
In nature, *magnesium* is found as *magnesite*, the carbonate  $\text{MgCO}_3$ , but far more commonly as *dolomite*, or magnesium limestone,  $\text{MgCO}_3 \cdot \text{CaCO}_3$ . Magnesium occurs in many natural silicates very widely distributed. *Hornblende* ( $\text{Mg}_2\text{CaFeSi}_4\text{O}_{12}$ ), *asbestos* ( $\text{Mg}_3\text{Si}_2\text{O}_7 \cdot 2\text{H}_2\text{O}$ ), *meerschaum* ( $\text{Mg}_2\text{Si}_3\text{O}_8 \cdot 4\text{H}_2\text{O}$ ), talcum, and many other complex silicates contain magnesium. All plants and animals contain magnesium compounds, and these finally appear in the ash as phosphates and carbonates. Magnesium ammonium phosphate occurs in guano and is useful as a fertilizer.

Metallic magnesium is sold in the market in powder, and in the form of wire and ribbon. It burns with a brilliant white light forming magnesium oxide,  $\text{MgO}$ . The metal is used for fireworks, for flash-lights in signaling and in photography.

Magnesium oxide and magnesium carbonate are used in medicine. The latter compound also is used as a face powder. Magnesium sulphate serves as a purgative under the name Epsom salt. All magnesium salts have a bitter taste and act as laxatives. When found in natural waters, the latter are called *bitter waters*. Some of them, like the waters of the springs at Epsom in England and Hunyad, in Hungary, are quite famous.

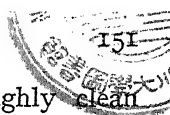
In their chemical behavior, magnesium compounds are in many ways very much like those of zinc.

*Zinc* is about 6.9 times as heavy as water, melts at  $420^\circ \text{C}$ ., and boils at  $918^\circ \text{C}$ . It is found in nature mainly as the carbonate  $\text{ZnCO}_3$  and the sulphide  $\text{ZnS}$ . These ores are first heated in contact with the air and are thus changed to zinc oxide. From its oxide, zinc is then obtained by heating with carbon in earthenware retorts.



In the form of sheets zinc is much used for gutters, roofs, ornaments, etc., on buildings. *Galvanized iron* consists of sheet iron which has been

## MAGNESIUM, ZINC, AND MERCURY



coated with zinc by the process of dipping the thoroughly clean iron into a bath of molten zinc. The beautiful flaky appearance of galvanized iron is due to the crystals of zinc on its surface. The coating protects the iron so that it will not rust. Zinc is also used in many primary electrical cells and in dry cells as the outer casing. There it serves as the negative pole of the battery.

Brass is the most important alloy of zinc which is in common use, but the latter metal is often present in other alloys. The soluble compounds of zinc are all poisonous.

*Zinc oxide* ( $\text{ZnO}$ ) is used in paints, being called zinc white or Chinese white. Mixed with lard or vaseline, it forms zinc oxide ointment, used in medicine. Zinc chloride,  $\text{ZnCl}_2$ , is very soluble in water. It is used for preserving wood, especially railway ties. The salt permeates the wood and is an antiseptic, killing the germs that are the cause of the decay of the wood. *White vitriol*, or zinc sulphate ( $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$ ), is another very common zinc salt. It is readily soluble in water.

In China, zinc is found in the southwest provinces, mainly, Kweichow and southern Hunan. Zinc blende, galena, and pyrites are the chief minerals found in the ore.

*Mercury*, *quicksilver*, hydrargyrum, is the only metal which is a liquid at ordinary temperatures. It is found in nature in the free state, but generally it is combined in *cinnabar* ( $\text{HgS}$ ), which is red in color and is used as a pigment, called vermilion. Mercury comes from Spain, Austria, Prussia, California, Japan, and China. In China there is a broad belt of it in northern Kweichow—running into Hunan—which has been worked for centuries. It is used in making lacquer.

Mercury is used in thermometers, barometers, in amalgams for the backs of mirrors, also for filling teeth. Its compounds are used in medicine. Mercurous chloride ( $\text{HgCl}$ ) is *calomel*. It is but slightly soluble in water. Mercuric chloride,  $\text{HgCl}_2$ , is *corrosive sublimate*. It is very soluble in water and is an exceedingly powerful poison. *All mercury compounds are poisonous*. The antidote is raw eggs and milk. The albumin forms an insoluble compound with the mercury salt, which is then gotten rid of by an emetic or purge.

Mercury melts at  $-39.4^\circ \text{C}$ . and boils at  $357^\circ \text{C}$ . Solid mercury may be hammered into sheets and cut with tools like other metals. The



*alloys* of mercury are called *amalgams*. The amalgam used in filling teeth commonly consists of tin, silver, and mercury. Gold alloys very readily with mercury. Indeed, the latter readily dissolves gold, and so is used in extracting fine particles of gold from the sands in which they occur. Mercury is also similarly used in silver mining.

#### QUESTIONS FOR STUDY

1. State the properties and uses of magnesium.
2. Name the chief ores of zinc. What is zinc blende?
3. Summarize the physical and chemical properties of zinc.
4. Mercuric nitrate and silver nitrate are both white soluble solids. How could you distinguish between them?
5. What properties make mercury useful in thermometers and barometers?

#### **Experiment 40. Identification of Simple Salts.**

**Apparatus.** Platinum wire or glass rod; charcoal or plaster block; blow-pipe; test tubes; Bunsen burner.

**Material.** Borax; solutions of cobalt nitrate, silver nitrate, barium chloride, and ferrous sulphate; dilute hydrochloric and nitric acids; concentrated sulphuric acid; unknown.

Use a very small portion of the unknown given you in making each test. Always keep a portion for proving your results.

*Keep a tabulated record of all tests, even those giving negative results.*

a. Determine the metallic part of the substance by means of (1) the flame test for sodium and potassium, (2) cobalt nitrate test, (3) borax bead test.

b. Then put a fresh portion of the unknown into a test tube and try to dissolve it in water, heating if necessary. Filter if there is an undissolved residue.

Divide the clear liquid into several portions and make tests for a chloride, a sulphate, and a nitrate.

c. Now determine whether the unknown contains a carbonate by the addition of an acid and making the limewater test of the gas that is given off.

Finally decide what you have found in the unknown substance, with reasons for your decision. Then take your notebook to the instructor.

# MAGNESIUM, ZINC, AND MERCURY

153

Number of unknown ————— .

TABLE

TEST	RESULT

The unknown contains ————— .

## CHAPTER XXXII

### IRON AND STEEL

Iron is doubtless the most valuable metal in the world. Not that it is so costly; indeed, its value rests upon its cheapness and its adaptability to an enormous number of uses. It has become a necessity in our modern life and we are said to live in *the age of steel and concrete*. For the development of the industry of iron and steel production chemical and engineering knowledge are required.

The principal compounds of iron which occur naturally are various forms of the oxides, the carbonate, and the sulphide. Of these the oxides are the chief source of metallic iron. The most important *iron ores* are hematite ( $\text{Fe}_2\text{O}_3$ ), magnetite ( $\text{Fe}_3\text{O}_4$ ), limonite ( $\text{Fe}_2\text{O}_3 \cdot 3\text{H}_2\text{O}$ ), and siderite ( $\text{FeCO}_3$ ).

Iron is the second in importance of China's mineral resources. Iron ore is found in every province but in most places only worked by native methods where coal is present. In Shansi, it is particularly abundant. The ores occur in the shales and sandstones of the carboniferous age. Owing to its scattered distribution there is not a sufficient supply of uniform ore to form a basis of blast furnace work on a large scale. There is much iron in western China—in Yünnan, Kweichow, and Szechwan.

Metallic iron is obtained from these oxides by the action of carbon and carbon monoxide at a high temperature. Iron ore is never pure iron compound but always contains certain amounts of other mineral matter, such as quartz or limestone. For this reason the smelting must be done in such a way that these impurities can be conveniently removed.

Iron differs from most of the other metals used in the industries in that the pure metal is seldom obtained and is of limited application, while that containing small percentages of other elements exhibits a wide variety of properties which make it of the greatest value for many different purposes.

Carbon is always present in amounts which vary from a mere trace to about 3 per cent. According to the condition of treatment, the carbon may be in the form of graphite scattered through the iron, or it may occur as a solid solution of carbon in iron, or as carbides of iron, one of the most important of which has the formula  $\text{Fe}_3\text{C}$ , and is called cementite. Manganese, silicon, and traces of phosphorus and sulphur, together with a little oxygen, are also present.

The properties of the iron are so much modified by the percentages of these elements, by their form of combination, and by the treatment of the metal during its production, that many varieties of iron are recognized in commerce, the chief of which are *cast iron*, *wrought iron*, and *steel*.

Ordinarily the first step in the manufacture of any variety of commercial iron is the production of cast iron. The ores are mixed with a suitable *flux* (material added to the charge which will combine with the impurities in the ore, to secure the removal of the *earthy matter* charged into the furnace along with the ore), and are reduced by heating with coke.

Cast iron varies considerably in composition, but always contains over 2 per cent of carbon, variable amounts of silicon, and at least traces of phosphorus and sulphur. The form in which the carbon is present, whether free or combined, also greatly modifies the properties of the iron. In general, cast iron is hard and brittle and melts at about  $1100^\circ$ . It cannot be welded or forged, but is easily cast in sand molds. It is rigid, but not elastic, and its tensile strength is small. It is used for making castings, but chiefly as a starting point in the manufacture of other varieties of iron.

**Blast Furnace Process.** The reduction of the iron ore is carried out in a large tower, called a blast furnace (see diagram). This is usually 80 ft. high and 20 ft. in internal diameter at its widest part, narrowing

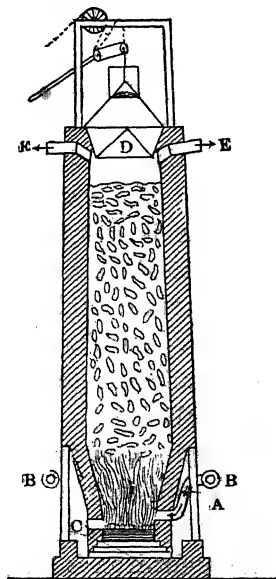


Fig. 63. Diagram of a Blast Furnace

somewhat toward both the top and the bottom. The walls are built of steel and are lined with fire brick. The base is provided with a number of pipes, A, called *tuyères*, through which hot air is forced into the furnace. The *tuyères* are supplied from a large pipe B, which girdles the furnace. At the base of the furnace is an opening, through which the liquid metal can be drawn off from time to time. There is also a second opening, C, somewhat above the first, through which the excess of slag overflows. The top is closed by a movable trap, D, called the *cone*, and through this the materials to be used are introduced. The gases resulting from the combustion of the fuel and the reduction of the ore, together with the nitrogen of the air admitted through the *tuyères*, escape through pipes, E. These gases are very hot and contain a sufficient percentage of carbon monoxide to render them combustible; they are accordingly utilized for heating the blast of air admitted through the *tuyères*, and as fuel for the engines.

Charges consisting of coke, ore, and flux in proper proportion are at intervals introduced into the furnace through the cone. The coke burns fiercely in the hot-air blast, forming carbon dioxide, which is at once reduced to carbon monoxide as it passes over the highly heated carbon.

Reduction of the ore begins at the top of the furnace through the action of the carbon monoxide. As the ore slowly descends, the reduction is completed, and the resulting iron melts and collects as a liquid in the bottom of the furnace, the lighter slag floating above it. After a considerable quantity of iron has collected, the slag is drawn off through C, and the iron is run into ladles and taken to the *converters* for the manufacture of steel; or it is run into sand molds and cast into ingots called *pigs*.

Wrought iron is made from cast iron by burning out most of the carbon, silicon, phosphorus, and sulphur, the operation being conducted in what is called a *puddling furnace*.

Wrought iron is soft, malleable, and ductile. Its tensile strength is greater than that of cast iron, but less than that of most steel. Its melting point is much higher than that of cast iron, and if melted, it is changed into steel. It is no longer produced to the same relative extent as in former years, since soft steel can be made at a less cost and has almost the same properties.

Steel, like wrought iron, is made from cast iron by burning out a part of the carbon, silicon, phosphorus, and sulphur, but the processes used are quite different from that employed in the manufacture of wrought iron. Nearly all the steel of commerce is made by one of two general methods known as the *acid Bessemer process* and the *basic open-hearth process*.

In the acid Bessemer process the furnaces used are lined with *silica*, which is an acid anhydride. These furnaces remove from the cast iron the carbon and silicon, but not the phosphorus and sulphur. The process is therefore employed when the cast iron to be used is low in phosphorus and sulphur.

*Details of Operation of Bessemer Process.* This process, invented about 1880, is carried out in great egg-shaped crucibles called *converters* (see diagram), each one of which will hold as much as 15 tons of steel. The converter is built of steel and lined with silica. It is mounted on trunnions, so that it can be tipped over on its side for filling and emptying. One of the trunnions is hollow, and a pipe connects it with an air chamber, A, which forms a false bottom to the converter. The true bottom is perforated, so that air can be forced in by an air blast admitted through the trunnion and the air chamber.

White hot liquid cast iron from a blast furnace is run into the converter through its open necklike top, B, the converter being tipped over to receive it; the air blast is then turned on, and the converter turned to a nearly vertical position. The carbon and silicon in the iron are rapidly oxidized (first the silicon, then the carbon), the oxidation being attended by a brilliant flame. The heat of reaction, largely due to the silicon, keeps the iron in a molten condition. The air blast is continued until the character of the flame shows that all the carbon has been burned away. The process requires from 15 to 20 minutes, and when it is complete, the desired quantity of carbon (generally in the form of high carbon iron alloy) is added, and allowed to mix thoroughly with the fluid. The converter is then tilted, and the steel run into molds, and the ingots so formed are hammered or rolled into rails or other objects.

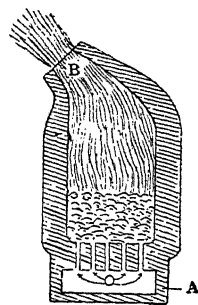


Fig. 64. Diagram of a Bessemer Converter

In the basic open-hearth process the lining of the furnace is made of limestone or dolomite, both of which act as bases. In such furnaces the phosphorus and sulphur are both removed as well as the silicon and carbon. The presence of more than traces of phosphorus and sulphur

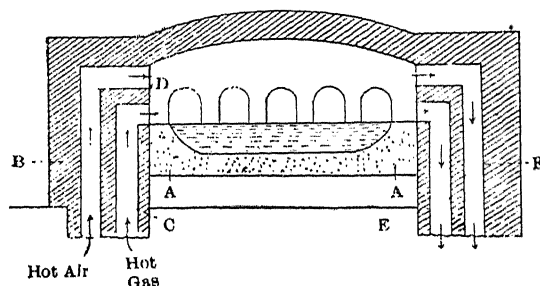


Fig. 65. Diagram of an Open-Hearth Furnace

in the finished steel renders the metal so brittle that it is worthless. The open-hearth process, therefore, possesses a great advantage over the Bessemer process in that it makes it possible to utilize iron ores (or cast iron obtained from them) that contain appreciable

quantities of phosphorus and sulphur. The operation does not need to be hastened, and steel of any desired composition can be produced.

*Details of the Open-Hearth Process.* Fig. 65 shows the simpler parts of the type of furnace used in this process. The hearth of the furnace is about 40 ft. in length, 12 ft. in width, and 2 ft. in depth, and is lined with limestone or dolomite (A, A). Either gas or sprayed oil is used as a fuel. Below the furnace is placed a checkerwork of brick so arranged that the hot products of combustion escaping from the furnace may be conducted through it, thus heating the bricks to a high temperature. Both the air necessary for combustion and the gaseous fuel (unless decomposed by heating, as in the case of natural gas and sprayed oil) are preheated by passing them over the hot bricks, so that the temperature reached during combustion is greatly increased.

The gas entering through C comes in contact at D with the hot air entering through B, and a vigorous combustion results, the flame passing above and over the cast iron and lime with which the furnace is charged. The products of combustion escape through E and F. At the temperature reached, the carbon in the cast iron is removed in the form of the oxide, the escaping gas giving the melted metal the appearance of boiling. The silicon, phosphorus, and sulphur unite with oxygen to form acid anhydrides; these combine with the lime to form a slag, and this rises to the surface of the melted charge and is easily removed.

When a test shows that the desired percentage of carbon is present, the melted steel is run into large ladles and then into molds. An average furnace produces about 50 tons of steel in a given charge, approximately eight hours being required in the process. At present by far the largest amount of steel produced in the United States is made by this process.

Steel contains from a trace up to 2 per cent of carbon, less than 0.1 per cent of silicon, and not more than traces of phosphorus and sulphur. When desired, a product containing as high as 99.85 per cent iron can be produced by the open-hearth method. Such steel is very soft, but resists rusting. As the percentage of carbon increases, the steel becomes harder and less ductile. Steel can be rolled into sheets, cast into molds, and forged into desired shapes.

When steel containing from 0.5 to 1.5 per cent of carbon is heated to a relatively high temperature and then cooled suddenly by plunging it into cold water or oil, it becomes very hard and brittle. When gradually reheated and then allowed to cool slowly, this hardened steel becomes softer and less brittle, and this process is known as *tempering*.

As we have seen, small quantities of carbon greatly modify the properties of iron, and equally marked effects may be produced by a great many other elements. Accordingly, to secure a steel with the requisite properties, suitable percentages of these elements are added to the steel just before it is run out of the furnace. The elements most frequently added are manganese, silicon, nickel, chromium, tungsten, vanadium, and titanium, and steel containing an appreciable percentage of any of these elements is called a *steel alloy*. The element is added in the form of a rich alloy of iron, such as ferrochromium or ferromanganese.

#### QUESTIONS FOR STUDY

1. Why does not iron occur in the native state? What does its native occurrence in meteors indicate?
2. Why is a blast furnace kept in continuous operation?
3. Compare the composition of cast iron, wrought iron, and steel.
4. Study topics—(a) production and transportation of iron ore in China, (b) uses of steel, (c) meteorites, (d) primitive iron smelting, (e) iron objects in the school building, (f) iron in the home.



**Experiment 41. Iron Salts in Photography—Blue Prints.**

**Apparatus.** 2 enameled pans, shallow, 8" × 10"; test tubes; graduate, 100 c. c.

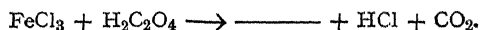
**Material.** Solution of ammonium ferric citrate (green scales), 100 g. in 1 liter; ferric chloride solution, 4 g. in 100 c. c.; potassium ferricyanide solution, 100 g. in 1 liter; oxalic acid solution, 50 g. in 1 liter; paper, unglazed, of good quality, cut into sheets 3" × 4".

*a.* Prepare two pieces of sensitive paper. This should be done in a darkened room, closet, or cupboard. Holding the paper by a corner, draw it lightly across the surface of a solution of ammonium ferric citrate contained in a shallow pan. Avoid getting the back of the paper wet with the solution, and make sure that the face is completely and evenly moistened. Hang the paper to dry in a dark place.

*b.* In a darkened room or cupboard, mix in a test tube 3 c. c. of ferric chloride with 3 c. c. of oxalic acid solution. Divide this mixture into two parts. Keep one part in the dark, while exposing the other to the direct rays of the sun for several minutes. To each of the tubes then add 3 c. c. of a solution of potassium ferricyanide.

*In which tube has a ferrous salt been produced?*

*Complete the equation:*



This is an example of a chemical action produced by light.

*Which of the substances may be regarded as having been reduced?*

*Which may be regarded as the reducing agent?*

*c.* Ammonium ferric citrate makes an excellent substance for preparing blue-print paper, because it contains within itself both the iron salt and the reducing agent.

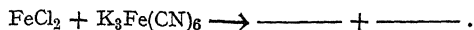
When the paper prepared with ammonium ferric citrate at the beginning of the period is thoroughly dry, arrange it for exposure to sunlight so that an opaque object, such as a key or a stencil cut in heavy paper, will keep the light from striking part of the paper.

The exposure should last three to ten minutes, depending on the strength of the light. At the end of this time, examine the paper quickly, and then immerse it, face down, in a pan containing a solution of potassium ferricyanide.

*Was any change apparent before the paper was put into the solution?*

*After it was put into the solution?*

*Complete the equation:*



Wash the paper thoroughly in running water, dry it, and paste in your notebook.

*What purpose does potassium ferricyanide serve in this operation?*

*What common photographic term might be applied to the potassium ferricyanide?*

*Why is it necessary to wash the paper thoroughly?*

*What photographic term can be applied to this part of the operation?*

## CHAPTER XXXIII

### IRON AND ITS COMPOUNDS

In small amounts iron is present almost everywhere. In all soils it occurs. To the brown sand, earth, and the brownish and yellowish clays it gives their characteristic colors. Sandstones and other rocks all contain compounds of iron. Indeed, the grains of sand in the sandstones are commonly cemented together with oxides of iron. Iron compounds are present in the tissues of all plants and animals. The green leaf contains chlorophyll, for the formation of which iron is essential, and the blood of animals contains hæmoglobin, in which iron plays an important rôle. In fact, without iron plants and animals cannot live. Nevertheless, it must be remembered that the quantity of iron present in living beings, important though it is, is after all, small; so, for instance, the human body contains only about 0.004 per cent of iron.

Iron forms two series of salts, the ferrous and the ferric. Thus there are *ferrous chloride* ( $\text{FeCl}_2$ ), and *ferric chloride* ( $\text{FeCl}_3$ ), *ferrous oxide* ( $\text{FeO}$ ), and *ferric oxide* ( $\text{Fe}_2\text{O}_3$ ). In the rusting of iron hydrated ferric oxide is formed. The black oxide, also known popularly as *hammer black*, which results when iron is heated in air, is *ferrous ferric oxide* ( $\text{FeO} \cdot \text{Fe}_2\text{O}_3$ , or  $\text{Fe}_3\text{O}_4$ ). It is magnetic. Iron nearly always acts as a basic element, and in this capacity it forms salts with practically all the various acids. Only a few of the most important salts of iron will be mentioned here.

*Ferrous sulphate*, also known as *green vitriol*, or *copperas*, is  $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$ . It is formed by the oxidation of *pyrite* ( $\text{FeS}_2$ ), which occurs in nature and is known as *fool's gold*, for its crystals are of a lustrous, golden yellow appearance. Ferrous sulphate dissolves readily in water. It is used in making ink. The latter contains tannin and ferrous sulphate. The *tannin* is an essential ingredient in extract of nutgalls, which is added to the ferrous sulphate solution. Ferrous sulphate is also used as a disinfectant, as a mordant in dyeing fabrics, and as a reducing agent. This salt is cheap and readily obtainable.

*Ferric chloride* ( $\text{FeCl}_3 \cdot 6\text{H}_2\text{O}$ ), forms a dark brown crystalline mass, not unlike maple sugar in appearance. It dissolves copiously in water, also in alcohol and ether. It is used in medicine. The so-called *styptic cotton*, which is used to stop the bleeding of wounds, consists of absorbent cotton treated with a solution of ferric chloride. This material not only allays bleeding by forming a clot with the blood, but it also acts as an antiseptic, thus protecting the wound from germs.

The change from ferrous to ferric compounds is called *oxidation*. The reverse change is called *reduction*. We have previously used these two words only when oxygen was involved in the reaction, but we shall now see that their meaning can be enlarged. When carbon monoxide burns it takes up one more atom of oxygen, *increasing its valence from two to four. This is an oxidation reaction.* Carbon is the positive element in carbon monoxide, and any *increase in the valence of a positive element is oxidation*; any *decrease* in the valence of such an element is *reduction*. It is not necessary that oxygen be the negative element which is added to the compound in order to increase its valence. In the case of iron chloride the addition of one chlorine atom to ferrous chloride produces ferric chloride; and this reaction is *exactly parallel to the addition of an oxygen atom to carbon monoxide; it is oxidation*. It will be well to bear in mind this new definition of oxidation and reduction.

#### QUESTIONS FOR STUDY

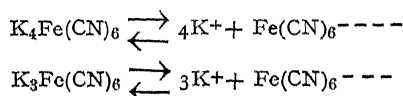
1. Why are the ferric compounds more stable than the ferrous?
2. Explain the fact that a solution of sodium carbonate precipitates the hydroxide and not the carbonate from a solution of a ferric salt.
3. What other oxidizing agents besides hydrogen peroxide might be used to change ferrous compounds to ferric?
4. How could you prove that a given solution contained *both* ferrous and ferric salts?

#### Experiment 42. Tests for Iron Salts.

**Apparatus.** Six test tubes; test-tube rack.

**Material.** Solutions of ferrous sulphate or other soluble ferrous salt (freshly prepared and thoroughly reduced), ferric chloride, potassium ferrocyanide, potassium ferricyanide (freshly prepared).

These tests involve the use of two substances that are complex cyanides. Potassium ferrocyanide ( $\text{K}_4\text{Fe}(\text{CN})_6$ ) is a compound of potassium cyanide ( $\text{KCN}$ ) and ferrous cyanide ( $\text{Fe}(\text{CN})_2$ ). Potassium ferricyanide ( $\text{K}_3\text{Fe}(\text{CN})_6$ ) is a compound of potassium cyanide ( $\text{KCN}$ ) and ferric cyanide ( $\text{Fe}(\text{CN})_3$ ). Note that the valence of iron in each of the compounds can be found by subtracting the number of potassium atoms from the number of cyanogen groups, and that the name of each of the compounds indicates the valence of the iron it contains. These two compounds ionize according to the following equations:



In writing the formulas of compounds that contain these complex ions it is sometimes desirable to use brackets, as in the formula for ferric ferrocyanide,  $\text{Fe}_4[\text{Fe}(\text{CN})_6]_3$ .

a. To 5 c. c. of a solution of ferrous sulphate (or other ferrous salt) add a few drops of potassium ferrocyanide solution.

*Record the color of the precipitate in the table.*

To 5 c. c. of a solution of ferric chloride (or other ferric salt) add a few drops of potassium ferrocyanide.

*Record the color of the precipitate in the table.*

This substance is known as Prussian blue.

*Write the equation for its formation.*

To a mixture of solutions of ferrous sulphate and ferric chloride add a few drops of potassium ferrocyanide solution.

*Which color obscures the other?*

*Is potassium ferrocyanide a test for the ferrous or for the ferric ion?*

b. To 5 c. c. of a solution of ferrous sulphate (or other ferrous salt), add a few drops of potassium ferricyanide.

*Record the color of the precipitate in the table.*

This is known as Turnbull's blue.

*Write the equation for its formation.*

To 5 c. c. of a solution of ferric chloride (or other ferric salt), add a few drops of potassium ferricyanide. Fill the test tube with water,

shake the contents, and examine the mixture closely, to see whether the liquid is clear or a precipitate is formed.

*Record the result in the table.*

To a mixture of solutions of ferrous sulphate and ferric chloride, add a few drops of potassium ferricyanide.

*Is the precipitate due to the ferrous or ferric salt?*

*Is potassium ferricyanide a test for the ferrous or for the ferric ion?*

TABLE

IRON SALTS	POTASSIUM FERROCYANIDE $K_4Fe(CN)_6$	POTASSIUM FERRICYANIDE $K_3Fe(CN)_6$
Ferrous Sulphate $FeSO_4$		
Ferric Chloride $FeCl_3$		

## CHAPTER XXXIV

### COPPER, GOLD, AND SILVER

The element copper has been used for various purposes since the earliest days of history. It is often found uncombined in nature, or combined with oxygen, sulphur, as the carbonate, or silicate. In Yünnan province, all the copper used in China and Burma was formerly produced.

Copper and gold are the only colored metals known. Copper melts at  $1084^{\circ}$ . It is tough, rather hard, ductile, and malleable. Much copper

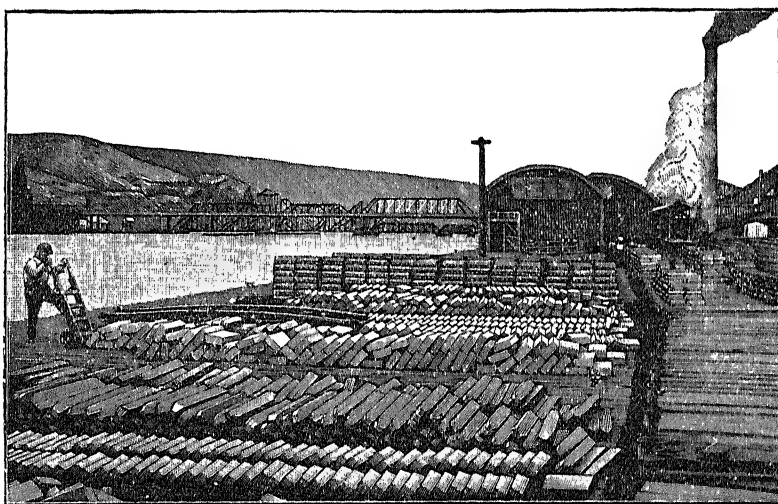


Fig. 66. Blocks of Copper Awaiting Shipment from Houghton, Michigan, U. S. A.

These ingots are made of ore from the Keweenaw Peninsula, "the copper country." One mining company alone has paid more than \$40,000,000 in dividends.

wire is used in electrical work, for it conducts heat and electricity well. For other purposes copper is generally used in the form of alloys, made by melting the metals together in the desired proportions. The following are some of the most important alloys of copper: *Brass*, which is yellow in color, contains one part zinc and two parts copper, though

other proportions are often used. *Dutch metal*, which is reddish brown, consists of one part zinc and five parts copper. *German silver* consists of 80 to 95 per cent brass and 5 to 10 per cent nickel. *Gun metal* contains nine parts copper in one part tin. The alloys of copper and tin are called *bronzes*. *Phosphor bronze* is used in making parts of machinery, as it is especially hard. Bronzes for statuary commonly contain three to eight parts tin, one to three parts lead, one to ten parts zinc, and eighty to ninety parts copper.

Copper dissolves readily in nitric acid, also in hot concentrated sulphuric acid, but cold sulphuric or hydrochloric acid has but very little effect on it. However, in the presence of air, many dilute acids very gradually attack copper, which fact must be kept in mind, for copper and brass are frequently used for cooking utensils, and *copper compounds are poisonous*. On copper roofs, coins, etc., that have been exposed to moist air for a long time there is formed a green deposit called *verdigris*. It is a basic carbonate of copper ( $\text{CuCO}_3 \cdot \text{Cu(OH)}_2$ ). In ammonia water, copper is soluble when in contact with the oxygen of the air. However, ammonia acts but slowly, hence it is often used for cleaning copper.

The most important as well as the most common salt of copper is copper sulphate, blue vitriol ( $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ ). It forms large blue crystals, is soluble in about three parts of water, and its solutions are used as a bath for *copper plating*, for spraying plants, especially in making Bordeaux mixture, and also for preparing other copper compounds. In Paris green, an insecticide, copper also is present.

Copper, like iron, forms two series of compounds; the *cuprous* compounds, in which it is univalent; and the *cupric* compounds, in which it is bivalent. The cupric salts are much the more common of the two.

*Gold* (aurum) has always been regarded as an article of value. It melts at  $1064^\circ$ , and is extremely malleable and ductile. In aqua regia it dissolves, forming gold chloride. *Gold coins* contain nine parts gold and one part copper. This alloy is much harder than pure gold, which is rather soft, and consequently will not wear well. For the same reason, copper alloys of gold are used for jewelry, ornaments, etc. Pure gold is twenty-four carats fine; sixteen-carat gold contains sixteen



parts gold and eight parts copper; fourteen-carat gold contains fourteen parts gold and ten parts copper, etc.

Gold salts are used in photography as "toning baths."

Many metallic articles are plated with gold. This is accomplished by immersing the article to be plated in a bath consisting of a solution of gold potassium cyanide ( $\text{KAu}(\text{CN})_4$ ), together with an electrode of pure gold; a current of electricity is then passed from the latter electrode through the solution to the article to be plated, which is readily coated with the metal. Gold is also used on the edges and backs of books, for ornamenting chinaware, gilding signs, etc.

*Silver* (argentum) occurs in nature in the free state, but also as the sulphide and chloride. It is much more abundant than gold. The metal melts at  $962^\circ$ , is malleable, ductile, and the best conductor of heat and electricity known. Sterling silver and silver coins contain one part copper and nine parts silver.

Nitric acid dissolves silver readily, forming silver nitrate ( $\text{AgNO}_3$ ), which is the most important of the silver compounds, for from it

all others are prepared. It is used in medicine and in photography. It is one of the most soluble of all salts. Even in ice-cold water it is possible to dissolve 122 parts in 100 parts of water, while in boiling water the solubility increases tenfold. This salt is also called lunar caustic and

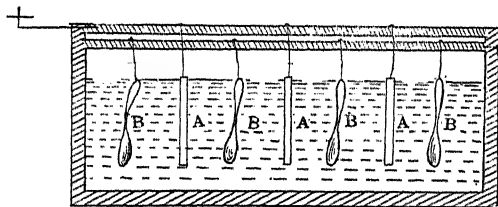


Fig. 67. The Process of Silver Plating

The anode A is made up of one or more plates of silver; the objects to be plated, B, is made the cathode.

is used for cauterizing wounds, etc.

Silver chloride, silver bromide, and silver iodide ( $\text{AgCl}$ ,  $\text{AgBr}$ , and  $\text{AgI}$ ) are all insoluble in water. These salts darken on exposure to light, probably because of separation of finely divided silver. This fact is the basis of the use of these salts in photography.

Silver plating is done in a solution of potassium silver cyanide ( $\text{KAg}(\text{CN})_2$ ). The process is similar in all respects to gold plating.

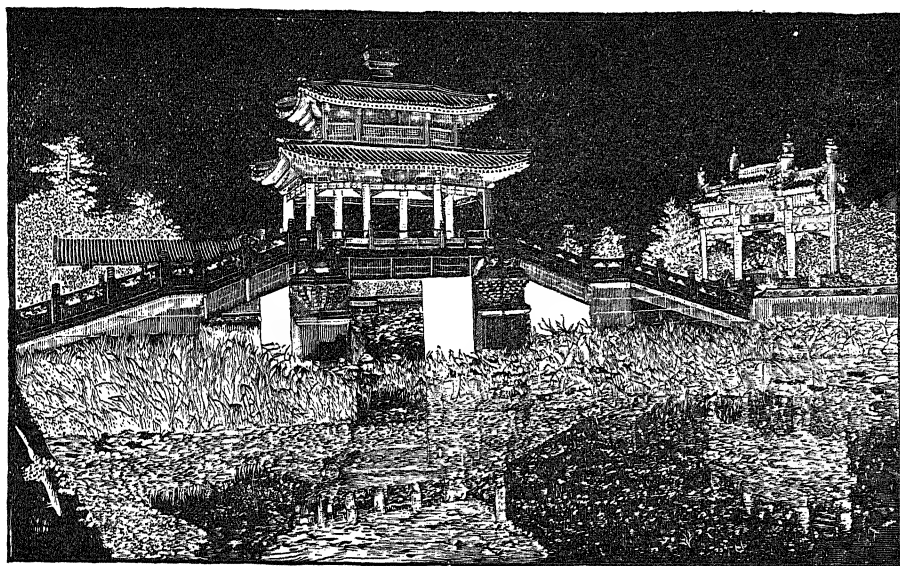


Fig. 68. Silver Salts in Photography — The Negative

The black stains that form on silverware (especially on spoons, forks, etc., which have been in contact with eggs) are silver sulphide, and not silver oxide as is often thought.

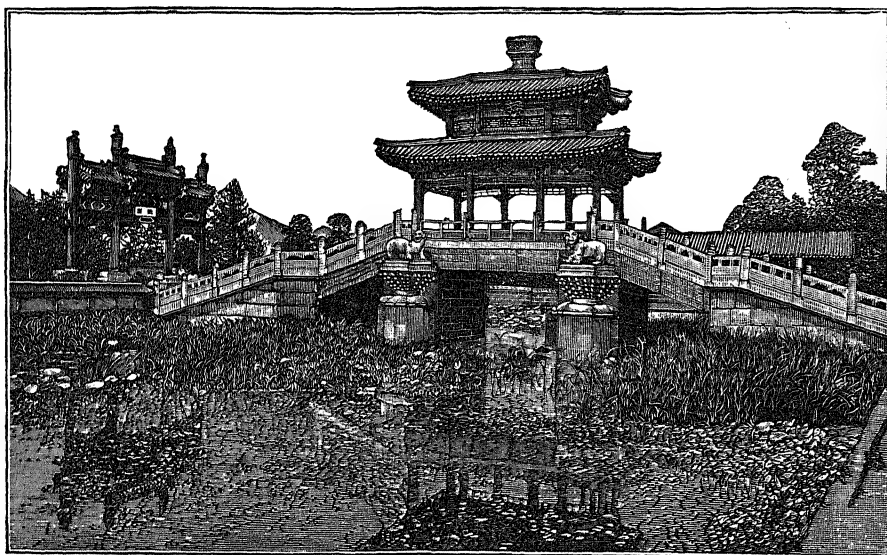


Fig. 69. The Positive Print

## QUESTIONS FOR STUDY

1. Describe the most common compound of copper. Of what use is it?
2. Mention three important alloys of copper.
3. For what purposes is gold used? Why?
4. What is the most important compound of silver? What are its uses?
5. What is sterling silver? For what is it used?

**Experiment 43. Preparation and Properties of Iodine.**

**Apparatus.** Seven test tubes; test tube rack; Bunsen burner; paper, 15 × 5 cm.

**Material.** Potassium iodide; manganese dioxide; sulphuric acid, 2 to 1; alcohol; potassium iodide solution; chloroform or carbon disulphide; chlorine water; bromine water; wooden splinters.

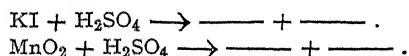
*a. Preparation.*

Mix on a piece of paper one gram of potassium iodide with one fourth its bulk of manganese dioxide. Roll the paper into a cylinder and insert it with its contents into a dry test tube held horizontally. Raise the tube to a vertical position, so that the mixture will fall to the bottom without touching the sides. Withdraw the paper, and add 2 c. c. of sulphuric acid (2 to 1). Warm the contents of the tube very gently.

*Results?*

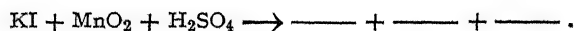
*Describe the deposit of iodine.*

*Complete the following equations, drawing lines through the formulas of substances that do not remain as permanent products of the reaction.*



*Write another equation to show how the two substances that you have crossed out react with each other.*

*Show the final products of the reaction by completing the following equation:*

*b. Effect of various solvents on iodine.*

With a splinter remove small portions of the solid that adheres to the sides of the tube, and try to dissolve the iodine in (1) water, (2) alcohol, (3) a water solution of potassium iodide, (4) carbon disulphide or chloroform.

*Record your results in the table.*

TABLE

SOLVENT	COLOR OF SOLUTION	DEGREE OF SOLUBILITY (SLIGHTLY, MODERATE, OR VERY)

*c. Carbon disulphide test for iodine.*

Put 3 c. c. of water and a very small crystal of iodine into a test tube; into another test tube put an equal volume of a solution of potassium iodide. To each test tube, add a few drops of carbon disulphide or chloroform. Shake each mixture thoroughly.

*In which case does the carbon disulphide acquire a color?*

*What is the color?*

*In what state must the iodine be to give the test with carbon disulphide?*

*d. Replacement of iodine by other halogens.*

To about 3 c. c. of a very dilute solution of potassium iodide, add a few drops of chlorine water.

*Result?*

Add chloroform or carbon disulphide and shake the mixture.

*Result?*

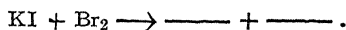
*Complete the equation:*



To another portion of potassium iodide solution, add bromine water and chloroform. Shake the mixture.

*Result?*

*Complete the equation:*



*Under what conditions does one element replace another? (See Experiment 18—Simple Replacement and table of Heats of Formation.)*

*Which of the halogens gives the greatest heat of formation in forming compounds with a given element?*

*Which gives least?*

*Which of the halogens is most easily replaced by other members of the family?*

*Which is least easily replaced?*

*Arrange the halogens in the order of their replacing power.*

## CHAPTER XXXV

### ALUMINIUM AND ITS COMPOUNDS

Although aluminium is the most abundant metallic element, it is found in nature only in the form of compounds. *Bauxite*, a hydrated oxide of aluminium, is the source of the metal and of most of its compounds. In Greenland a double fluoride of aluminium and sodium is

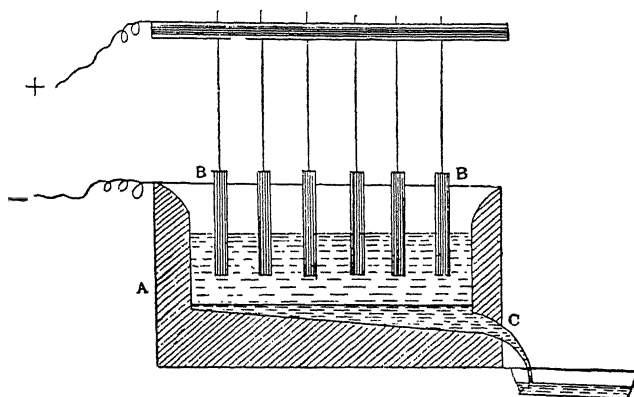


Fig. 70. Diagram Illustrating the Manufacture of Aluminium

An iron box A, about 8 ft. long and 6 ft. wide, is connected with a powerful electrical generator in such a way as to serve as the cathode upon which the aluminium is deposited. Three or four rows of carbon rods, B, B, dip into the box and serve as the anodes. The box is partly filled with cryolite, and the current is turned on, generating enough heat to melt the cryolite, aluminium oxide is then added, and acts as an electrolyte, being decomposed into aluminium and oxygen. The temperature is maintained above the melting point of aluminium, and the liquid metal, being heavier than the cryolite, sinks to the bottom of the vessel, from which it is tapped off from time to time through the tap hole C.

found which is called *cryolite* ( $\text{AlF}_3 \cdot 3\text{NaF}$ ). Aluminium is also found as a constituent of all the common siliceous rocks, such as feldspar, clay, mica, and slate. *Emery*, a very useful and valuable natural abrasive is an impure form of aluminium oxide ( $\text{Al}_2\text{O}_3$ ), colored brown by iron oxide. *Sapphires* and *rubies*, which are highly prized as gems, are beautifully crystallized aluminium oxide.

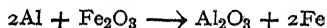
Because of the very great chemical activity of the element, aluminium oxide cannot be reduced to metallic aluminium by means of carbon in the furnace. The metal is now obtained by the *electrolysis* of its oxide.

Aluminium is a silver-white metal which usually has a rather dull surface due to a thin coating of the oxide. It is very light in weight, being only one third as heavy as iron, and is an excellent conductor of both heat and electricity. Hydrochloric acid acts on it rather rapidly, producing hydrogen. It is dissolved by sodium and potassium hydroxide with the formation of compounds called *aluminates* and with the liberation of hydrogen.

Aluminium is not so cheap as steel, and its use is therefore restricted to articles in which lightness is a prime essential, such as parts of automobiles and aëroplanes. It is sometimes used in the place of copper as a conductor of electricity. It is a very satisfactory material for cooking utensils, not only because of its lightness, but because it keeps bright, since the thin film of oxide is transparent and protects the metal from further action of the air. Powdered aluminium is used extensively as a paint to protect other metals from corrosion. Very thin foil of aluminium is replacing tin foil to some extent.

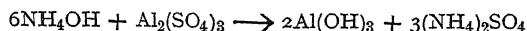
There are several important alloys of aluminium. Copper containing 5 to 12 per cent of aluminium is called *aluminium bronze*; it has a brilliant yellow-gold color. *Magnalium* contains a small amount of metallic magnesium. These alloys can be worked more easily than the pure metal.

Metallic aluminium is a very powerful reducing agent. It reacts with the oxides of many metals, producing aluminium oxide and the free metal. This reaction takes place with the evolution of much heat and when once started proceeds vigorously. A mixture of aluminium and iron oxide is called *thermite*. When this mixture is ignited by means of a burning strip of magnesium ribbon, a reaction takes place, and a temperature of 3000° to 3500° is reached.

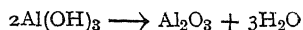


Advantage is taken of this reaction in *welding* rails, shafting, and similar objects.

The hydroxide is formed as a colloidal precipitate when ammonium hydroxide is added to a salt of aluminium.



When heated, the aluminium hydroxide is decomposed into the oxide and water.



Aluminium hydroxide has both *acid* and *basic* properties. It dissolves in most acids to form soluble salts, and in the strong bases to form aluminates.

When a solution of aluminium sulphate is mixed with a solution of potassium sulphate and evaporated, a double salt crystallizes out which has been long known as *alum*  $[\text{K}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}]$ . There are a number of double salts which have a similar formula; such as ammonium alum, and chrome alum. Ordinary alum is the commonest form in which aluminium compounds are used.

When aluminium hydroxide is precipitated in muddy water, it has the power of carrying down with it suspended particles. This property is used to advantage in the purification of water. Alum and lime are added in proper proportions to the contaminated water supply; the aluminium hydroxide which is formed carries down the suspended material from the water.

Clay consists essentially of aluminium silicates that have been formed from the weathering of feldspars and other complex silicates that make up the granitic rocks. In its purest form clay is white and is called *kaolin*  $[\text{H}_2\text{Al}_2(\text{SiO}_4)_2 \cdot 2\text{H}_2\text{O}]$ . This is used for making white porcelain ware. Common clay is discolored by impurities. Usually it is brown or red which is caused by the presence of oxides of iron. Ordinary clay serves for making bricks, flowerpots, and the inferior and cheaper grades of pottery and crockery ware. *Porcelain* is vitreous throughout. This is accomplished by mixing finely pulverized feldspar and quartz with the kaolin before molding the dishes. When these are dried and finally fired, the feldspar and quartz fuse and fill the pores of the ware so that it exhibits a perfectly vitreous surface instead of an earthy one, as the cheaper earthenware, either glazed or unglazed, always does.



Fire bricks contain a larger amount of silica than ordinary ones and are consequently more refractory. Fire clay is especially rich in silica. Colored porcelain and colored glazes are produced by means of various metallic oxides, as in making colored glass.

*Aluminium sulphate*  $[\text{Al}_2(\text{SO}_4)_3]$  is made by dissolving aluminium hydroxide in sulphuric acid. This salt is readily soluble in water and is used as a mordant for fixing dyestuffs upon fabrics. In the paper industry it serves in sizing paper.

#### QUESTIONS FOR STUDY

1. Where is aluminium found in nature, and in what form?
2. Mention the most important characteristics of metallic aluminium. For what is it used?
3. What is thermite? For what is it used?
4. What is the difference between porcelain and ordinary crockery ware?

#### **Experiment 44. Aluminium Hydroxide.**

**Apparatus.** Hydrometer jar, or cylindrical graduate, 250 c. c.; three test tubes; glass stirring rod, 15".

**Material.** Aluminium sulphate, 20 g. to the liter; limewater; fine clay; ammonium hydroxide solution, 1 to 10; logwood solution; alizarine mixed with water.

##### *a. Preparation.*

To one sixth of a test tube of a dilute solution of aluminium sulphate  $[\text{Al}_2(\text{SO}_4)_3]$ , add twice as great a volume of limewater.

*Result?*

*What is the appearance of this precipitate of aluminium hydroxide?*

*Write the equation for the precipitation.*

##### *b. As a coagulum.*

Add water to a cylindrical jar until it is about two thirds full. Render the water turbid by stirring in a little fine clay. Then pour in one third of a test tube of aluminium sulphate solution and mix it with the muddy water by thorough stirring.

To this mixture, add slowly, without stirring, two thirds of a test tube of limewater. Allow the water to stand.

Note what is happening from time to time.

*Result?*

*What precipitate was formed in the muddy water?*

*How are the suspended particles of clay removed so as to leave the water clear?*

*c. Formation of lakes.*

To one sixth of a test tube of logwood solution, add ammonium hydroxide. Note that no precipitate is formed, although the logwood changes somewhat in color.

To a solution of aluminium sulphate in another test tube, add some ammonium hydroxide.

*What is the precipitate?*

Add some logwood solution to the test tube containing the precipitate. Shake the test tube and allow the contents to settle.

*What does the color of the liquid in the test tube show about the amount of logwood in solution?*

*Compare the color of the precipitate with the color of logwood and with the color of aluminium hydroxide.*

Such a combination of a dye with aluminium hydroxide or other suitable compound, is commonly called a "lake."

Formerly, aluminium hydroxide was extensively used in dyeing cloth. Its gelatinous nature enables it to adhere to the fibers of the cloth. It also adheres to the dye. Thus it acts as a binder between the cloth and the dye. A substance that has this characteristic is said to be a *mordant*.

*State how you would dye a piece of white cotton cloth by using the same three solutions employed to make the logwood lake.*

Using aluminium sulphate solution, ammonium hydroxide solution, and alizarine, make an alizarine lake.

*How does the alizarine lake differ in color from the original alizarine and from the aluminium hydroxide?*

## CHAPTER XXXVI.

### TIN AND LEAD

Tin is found in nature chiefly as the oxide  $\text{SnO}_2$ , called cassiterite, or tinstone. Practically all the world's supply comes from Cornwall, the East Indies, and Bolivia.

The metallurgy of tin is very simple. The ore, separated as far as possible from earthy matters, is mixed with carbon and heated in a furnace, the reduction taking place readily.

Pure tin, called *block tin*, is a soft white metal with a silverlike appearance; it melts readily ( $232^\circ$ ) and is somewhat lighter than copper, having a density of 7.3. It is malleable and can be rolled out into very thin sheets, forming *tin foil*; most tin foil, however, contains a marked percentage of lead.

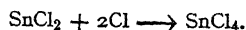
Under ordinary conditions tin is unchanged by air or moisture, but at a high temperature it burns, forming the oxide  $\text{SnO}_2$ . Dilute acids have little effect upon it, but concentrated acids attack it readily.

A great deal of tin is used in making *tin plate*. The process consists in dipping thin sheets of iron into the melted tin and is quite similar to that of galvanizing iron. Owing to its resistance to the action of air and weak acids, tin plate is used in many ways, such as in roofing, in the manufacture of tin cans, cooking vessels, etc. Small pipes of block tin are used instead of lead for conveying pure water or liquids containing dilute acids, such as soda water. Many useful alloys contain tin. *Pewter* and soft solder are alloys of tin and lead.

Tin forms two series of metallic compounds: the *stannous*, in which the tin is bivalent, as is illustrated in the compounds  $\text{SnO}$ ,  $\text{SnS}$ ,  $\text{SnCl}_2$ ; and the *stannic*, in which it is tetravalent, as shown in the compounds  $\text{SnO}_2$ ,  $\text{SnS}_2$ .

*Stannous* chloride is prepared by dissolving tin in concentrated hydrochloric acid and evaporating the solution to crystallization. The

crystals which are obtained have the composition  $\text{SnCl}_2 \cdot 2\text{H}_2\text{O}$ , and are known as *tin crystals*. By treating a solution of stannous chloride with aqua regia, *stannic* chloride is formed:



*Lead* is found in nature chiefly as the sulphide  $\text{PbS}$ , called galena. In this country the best-known mines are in Hunan. Lead and zinc ores are produced there with their by-products.

Almost *all* the lead of commerce is made from galena, which usually contains some silver. To obtain this silver most economically, it is customary to combine richer silver ores with lead ores and work the two together.

Lead is a heavy metal which has a brilliant silvery luster on a freshly cut surface, but which soon tarnishes by oxidation to a dull blue-gray color. It is soft, easily fused (melting at  $327^\circ$ ), and malleable, but has little toughness or strength.

It is not acted upon to any great extent, under ordinary conditions, by the oxygen of the air, but at a high temperature is changed into the oxide. With the exception of hydrochloric and sulphuric acid (which form insoluble compounds) most acids, even very weak ones, act upon it, forming soluble lead salts. Hot concentrated hydrochloric and sulphuric acid also attack it to a slight extent.

Lead finds many important applications in the industries, chiefly in the manufacture of storage batteries, in linings for sulphuric acid plants, in alloys of various kinds (such as shot, antifriction metals, type metal, and pewter), and in water pipes for plumbing. Since lead dissolves to some extent in pure water, it should not be used for pipes that are to carry rain water. About one third of the production of lead is used in making paint, and is permanently lost.

In nearly all its compounds, lead is bivalent, but in a few of its compounds it has a valence of four. All its compounds are poisonous. Lead forms a number of oxides, the most important of which are—*litharge* ( $\text{PbO}$ ), with a number of commercial uses; *minium*, or *red lead* ( $\text{Pb}_3\text{O}_4$ ), used as a pigment, or as an ingredient of a cement used in joining gas pipes; and *lead peroxide* ( $\text{PbO}_2$ ), which is a good oxidizing agent.

While the normal carbonate of lead,  $\text{PbCO}_3$ , is found to some extent in nature and can be prepared in the laboratory, basic carbonates of varying composition are much easier to obtain. One of the simplest of these has the composition  $(\text{PbCO}_3)_2 \cdot \text{Pb(OH)}_2$ , and is called *white lead*. This is prepared on a large scale as a white pigment and as a body for paints which are to be colored with other substances.

#### QUESTIONS FOR STUDY

1. How could you detect lead if it were present in tinfoil?
2. Stannous chloride reduces gold chloride ( $\text{AuCl}_3$ ) to gold. Give equation.
3. What sulphates other than lead are insoluble?
4. What weight of tin could be obtained by the reduction of 1,000 kilograms of cassiterite?
5. What reaction would you expect to take place when lead peroxide is treated with hydrochloric acid?

#### Experiment 45. Qualitative Separation of Lead, Silver, and Mercury.

**Apparatus.** Test tubes; test-tube racks; funnel; Bunsen burner.

**Material.** Hydrochloric acid, concentrated and dilute, nitric acid, concentrated and dilute; ammonium hydroxide, 1 to 3; solutions of lead, silver, and mercurous nitrates; potassium chromate solution; copper strip; filter paper; unknown solutions.

a. In one test tube take 10 c. c. of a solution of lead nitrate; and in another, 10 c. c. of a solution of silver nitrate. To both test tubes add dilute hydrochloric acid till the reaction is complete.

*Results?*

*Write the equations.*

Allow the precipitates to settle and then pour off top liquid from each of the two test tubes. Add to the precipitates in the test tubes enough cold water to nearly fill the tubes, and shake the contents. Again let the precipitates settle and then pour off the top liquid.

*What compound has been removed by washing the precipitate and then pouring off the top liquid?*

Try the effect of hot water on the precipitate of lead chloride.

*Result?*

Divide the precipitate of silver chloride between two test tubes. With one part try the effect of hot water; with the other, the effect of ammonium hydroxide.

*Results?*

b. To 10 c. c. of a solution of mercurous nitrate ( $\text{HgNO}_3$ ), add dilute hydrochloric acid till the action is complete.

*Result?*

*Write the equation.*

Wash the precipitate with cold water and divide it between two test tubes. Find out whether hot water dissolves the mercurous chloride.

*Result?*

*What effect does ammonium hydroxide have upon the mercurous chloride?*

c. In the same test tube take 5 c. c. each of solutions of silver, lead, and mercurous nitrates. Add dilute hydrochloric acid till precipitation is complete.

*Of what does the precipitate consist?*

Filter. Wash the precipitate on the filter paper with a very little cold water. Next wash the precipitate thoroughly with hot water, keeping the washings.

*Which one of the chlorides was dissolved by hot water?*

To confirm this, add to the hot filtrate a solution of potassium chromate,  $\text{K}_2\text{CrO}_4$ . This chromate gives an insoluble and characteristic compound of the metal whose chloride is soluble in hot water.

*Write the equation for this confirmatory test.*

*Give name, formula, and color of the characteristic compound formed.*

Wash the precipitate remaining on the filter paper with ammonium hydroxide, keeping the washings.

*Which chloride gives the color?*

*What chloride is contained in the ammonium hydroxide filtrate?*

Prove the presence of this chloride by adding a slight excess of nitric acid.

*Name the precipitate and state the characteristic properties by which you recognize it.*

To dissolve the precipitate still remaining on the filter paper, add a little aqua regia (nine drops concentrated hydrochloric acid to three

drops concentrated nitric acid). Dilute with water the solution thus obtained, and put into it a bright strip of copper. After several minutes, remove the strip, and wash and rub it.

*Result?*

*Explain why this dissolving in aqua regia and the addition of a copper strip is a confirmatory test.*

d. Obtain from the instructor an unknown solution. Using the methods in c, analyze the solution for lead, silver, and mercury.

*Record all the steps, even those giving negative results.*

*Underline the metal found in your unknown.*

lead

mercury

silver





hence it is often used to plate iron, copper, brass, etc. *Nickel plating* is carried on in the same way as gold or silver plating. The bath for nickel plating consists of a solution of nickel ammonium sulphate  $[(\text{NH}_4)_2\text{SO}_4 \cdot \text{NiSO}_4 \cdot 6\text{H}_2\text{O}]$ , and, of course, a plate of solid nickel is placed in the solution opposite to the object to be plated. The electric current is then passed from the nickel plate through the solution to the object that is to receive the nickel coating. The salts of nickel are green, and when soluble they yield green solutions.

*Cobalt* is analogous to nickel in most important respects. Its salts when dissolved in water yield dark red solutions, and the crystals that deposit from such solutions are also red, for in general the crystals contain water of crystallization. When this water is driven off, anhydrous, i. e., dry compounds are formed which are blue. Cobalt silicate is blue, and so glass containing cobalt silicate is blue glass, also called *smalt glass*. Indeed, about the only practical use which is at present made of cobalt is in the production of blue glass, blue porcelain and blue glazes on enamel ware. Cobalt, nickel, and iron are metals that are attracted by a magnet, and are consequently said to be *magnetic*.

#### QUESTIONS FOR STUDY

1. What is manganese dioxide used for? Write the equation expressing the chemical changes.
2. Of what use is potassium permanganate?
3. What use is made of potassium bichromate?
4. State the characteristics and uses of nickel.
5. What use is made of cobalt?

#### **Experiment 46. Soap Making.**

**Apparatus.** Bunsen burner; ring stand; teaspoon; evaporating dish; graduate, 25 c. c.; stirring rod.

**Material.** Beef fat (rendered by heating it in an evaporating dish until the fat is melted, and then pouring off the fat from the connective tissue), or lard; 33 per cent solution of sodium hydroxide.

Place in an evaporating dish two level teaspoonfuls of "rendered" beef fat or lard. Heat the dish with a small flame until the fat is melted. Then add 3 c. c. of a 33 per cent solution of sodium hydroxide.

Warm the dish very gently with a small flame, stirring the contents constantly. Continue the heating until there is a soapy mass in the dish. Then allow the dish to cool. The residue left in the dish is soap.

Beef fat is mainly glyceryl stearate  $[C_3H_5(C_{18}H_{35}O_2)_3]$ . Stearic acid has the formula  $H(C_{18}H_{35}O_2)$ .

*Write the equation for the reaction between the glyceryl stearate and the sodium hydroxide, producing sodium stearate (a soap) and glycerine  $[C_3H_5(OH)_3]$ .*

Organic fats and oils are esters of the alcohol glycerine and fatty acids.

*How are such fats and oils converted into soap?*

*What alcohol is a by-product of soap making?*

*What became of it in the case of the soap just made?*

## CHAPTER XXXVIII

### CARBON COMPOUNDS

The most striking characteristic of carbon is its ability to form almost numberless compounds in which a whole series of carbon atoms are linked to one another. These are called *organic compounds*; compounds of all the other elements are called *inorganic compounds*. We can define organic chemistry as *that* branch of chemistry which deals with the compounds of carbon. The elements which often appear with carbon are hydrogen, oxygen, nitrogen, the halogens, and sulphur; a few other elements sometimes are constituents of organic compounds, but they are of minor importance.

*Hydrocarbons*, which are compounds of carbon and hydrogen, are very numerous, and they may be regarded as the mother substances from which all other compounds of carbon are derived. The simplest hydrocarbon is *methane*, or marsh gas ( $\text{CH}_4$ ). From marsh gas a variety of products may be obtained. Thus by treatment with chlorine there may be obtained, successively, methyl chloride ( $\text{CH}_3\text{Cl}$ ), methylene chloride ( $\text{CH}_2\text{Cl}_2$ ), chloroform ( $\text{CHCl}_3$ ), and carbon tetrachloride ( $\text{CCl}_4$ ). Of these compounds chloroform and carbon tetrachloride are by far the most important ones. Both are liquids at ordinary temperatures. Chloroform boils at  $61^\circ \text{C}$ ., is heavier than water, has a rather agreeable odor, and is used as an anæsthetic. The analogous iodine compound, *iodoform* ( $\text{CHI}_3$ ), is a yellow crystalline solid of rather disagreeable odor. It is used in surgery as a dressing for wounds, especially when pus has formed. Carbon tetrachloride is also a colorless, heavy liquid which does not dissolve in water. It boils at  $76^\circ$  and is not inflammable. It is an excellent solvent for fats and hence is often used for removing grease spots from clothes. It is also used in fire extinguishers. Carbon tetrachloride extinguishes fires because it evaporates and crowds the air away from the burning substances, thus making it impossible for combustion to proceed further.



really an alcohol. Its composition is expressed by the formula  $C_3H_5(OH)_3$ . All alcohols contain the OH group. They may all act as bases, reacting with acids to form organic salts and water.



It will be seen at once that this reaction is similar to that when sodium hydroxide acts on acetic acid:



Organic salts, like ethyl acetate, in which there is a hydrocarbon radical which plays the rôle of a metal in that it has basic properties, are called *etheral salts*, or *esters*. Esters are quite common in nature. So, for example, *wintergreen oil*, methyl salicylate, is an ester of salicylic acid in which methyl ( $CH_3$ ), the radical from wood alcohol, acts as a base. *Banana oil*, amyl acetate ( $CH_3COO \cdot C_5H_{11}$ ), is formed when acetic acid and amyl alcohol react with each other.

The fats are esters in which stearic, palmitic, and oleic acids act as acids and glycerine as the base. So in mutton tallow we have principally *glyceryl stearate*, in palm oil chiefly *glyceryl palmitate*, while *hog's lard*, *cotton seed*, and *olive oil* are rich in *glyceryl oleate*. In general, the fats are mixtures of these three salts of glycerine. The softer the fat, the richer it is in oleine or glyceryl oleate; the harder the fats, the richer they are in stearine or glyceryl stearate. On heating the fats with caustic soda or caustic potash the corresponding salts of sodium or potassium are formed and glycerine is simultaneously produced. These sodium salts of stearic, palmitic, and oleic acids are the *hard soaps*. The corresponding potassium salts are the *soft soaps*.

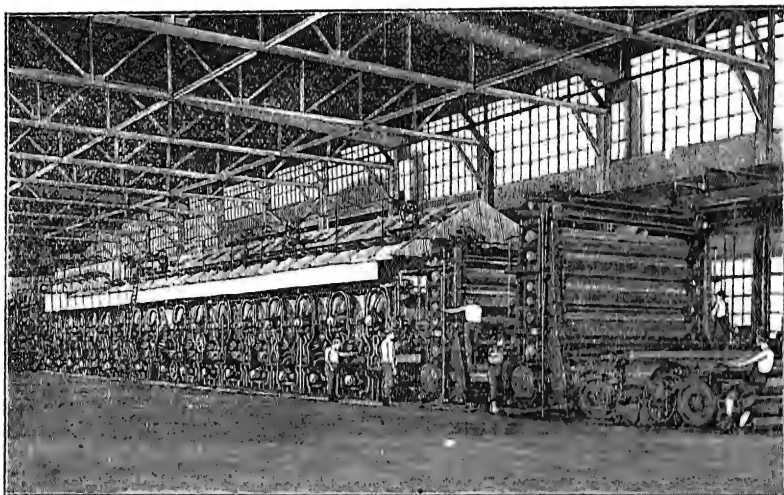
From the hydrocarbon benzene  $C_6H_6$ , are derived many other useful substances. *Phenol*, or carbolic acid ( $C_6H_5OH$ ), is used as an antiseptic. *Benzene* and *toluene* ( $C_7H_8$ ) are liquids; *naphthalene* ( $C_{10}H_8$ ) and *anthracene* ( $C_{14}H_{10}$ ) are solids. These hydrocarbons are indispensable, being the present substances of dyes, medicines, and explosives.

Among the many other *organic acids* that exist there are the following which are frequently met in daily life: (1) *Oxalic acid*  $[(COOH)_2]$ , which is made from sawdust by oxidizing the latter with the aid of nitric acid. It serves for removing ink and rust spots from floors and fabrics, and is



mineral content, they consist of almost pure cellulose. The composition of cellulose is expressed by the formula  $(C_6H_{10}O_5)_x$ . In the form of hay, straw, etc., cellulose serves as a food for cattle, horses, and other herbivorous animals; for fuel, clothing, and shelter, cellulose in various forms is invaluable to mankind. *Paper* is made from rags, straw, wood pulp, etc.; it therefore consists of fibers of cellulose that are matted together.

*Nitrates of cellulose* may be formed by treating cellulose with a mixture of nitric and sulphuric acids. These *nitrocelluloses* are true esters of cellulose, for on saponifying them with caustic soda, sodium nitrate and cellulose are formed. *Guncotton* is cellulose hexanitrate  $[C_{12}H_{14}-$



From the *National Geographic*

Fig. 72. In Five Minutes This Plant Can Turn Out a Mile of Paper

The Kalamazoo Valley is among the most extensive paper manufacturing centers of the world. This is one of its largest units and embodies the most advanced ideas in paper-making machinery.

$(NO_2)_6O_4]$ . It can be made to explode violently by means of a fulminating cap of mercury, and serves as an explosive. It may be used alone or with *nitroglycerine* which is produced when glycerine is acted upon by a mixture of concentrated nitric and sulphuric acids. When absorbed in diatomic earth and molded in sticks it forms *dynamite*. Guncotton, nitroglycerine, and dynamite are used in the arts for





The *sugars* may be divided into (a) *glucose*, or *dextrose*, also known as *grape sugar*, (b) *levulose* or *fructose*, or *fruit sugar*. The composition of the fructoses corresponds to the formula  $C_{12}H_{22}O_{11}$ , and the chief common representatives of this class are (1) *cane sugar*, also called *saccharose* or *sucrose*; (2) *maltose*, or *malt sugar*; (3) *lactose*, or *milk sugar*.

*Glucose* is found in the juice of many fruits, especially of grapes, whence the name *grape sugar*. Commercially, it is prepared on a large scale by heating starch with dilute sulphuric acid. *Glucose* is a good, wholesome food. It is mainly used in candies, sirups, etc. It is about three fifths as sweet as cane sugar.

*Levulose*, or *fructose*, is also found in the juice of fruits. It also occurs in honey. Cane sugar is obtained commercially from sugar cane and sugar beets. In sorghum, in maple sap, in nuts, and in the blossoms of many plants *saccharose* abounds also. *Maltose*, or *malt sugar*, occurs in malt. It is an intermediate product in the formation of alcohol from starch, for when *diastase*, an enzyme contained in malt, acts upon starch, *maltose* is produced. This then may be fermented to alcohol and carbon dioxide. Upon these facts the commercial production of fermented liquors like beer depends. *Lactose*, or *milk sugar*, occurs in the milk of mammals. Cow's milk contains about 5 per cent of *lactose*. It is not so sweet as cane sugar.

The *proteins* consist of carbon, hydrogen, oxygen, nitrogen, and sulphur. They are an exceedingly important class of substances, for when the fats, mineral matter, and water are taken from the body of an animal what is left consists entirely of proteins. No animal can live without proteins as food. The *nucleoproteins* often contain 5 to 6 per cent of phosphorus. *Albumins* are proteins, and occur in eggs, in muscles, in milk, in blood serum, in the seeds of plants especially in peas, beans, and other legumes.

#### QUESTIONS FOR STUDY

1. What is the difference between a hydrocarbon and a carbohydrate? Name 5 examples of each.
2. What is carbon tetrachloride? What are its uses?
3. How is ordinary alcohol produced? What are its uses?

4. To what class of substances do the plant and animal fats and oils belong?
5. Mention 5 organic acids and tell where they occur.
6. To what animals may cellulose serve as a food? Will cellulose in any form do for this purpose? Explain by a few examples.
7. Classify the sugars.
8. What is a protein? Give several examples.

**Experiment 47. Preparation of Ethereal Salts (Esters).**

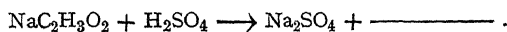
**Apparatus.** Bunsen burner; three test tubes; notched cork to fit test tube.

**Material.** Sulphuric acid, concentrated; sodium acetate; salicylic acid; ethyl alcohol, 95 per cent; amyl alcohol; methyl alcohol.

*a. Ethyl acetate.*

Dissolve enough sodium acetate to fill the curved bottom of a test tube in a very little water, and then add a few drops of concentrated sulphuric acid. Acetic acid is produced,  $\text{H}(\text{C}_2\text{H}_3\text{O}_2)$ .

*Complete the equation:*



To the contents of the test tube add a few drops of ethyl alcohol. Warm the tube gently, and note a distinctive odor, different from that of alcohol. This is due to ethyl acetate ( $\text{C}_2\text{H}_5\cdot\text{C}_2\text{H}_3\text{O}_2$ ).

*Write the equation for its formation.*

*b. Amyl acetate.*

Put one drop of amyl alcohol ( $\text{C}_5\text{H}_{11}\text{OH}$ ) into a test tube, add two drops of concentrated sulphuric acid, and then add a small pinch of sodium acetate. Warm the test tube gently. Remove the tube from the flame, and, from a test tube with a notched cork, let water run, a drop at a time, down the side of the tube, until about 3 c. c. have been added. Again warm the test tube, shaking it as you do so. Note the distinctive odor. Amyl acetate has been formed.

*In what fruit have you noted this odor?*

*c. Methyl salicylate.*

Put a pinch of salicylic acid in a test tube. Add a drop of methyl alcohol, and about twenty drops of concentrated sulphuric acid. Warm the mixture *very gently*, shaking the tube as you do so. Note the odor produced.

*Of what familiar substance does it remind you?*

*d. General questions.*

*What are the products of the reaction between a base and an acid?*

*What radical is common to bases and alcohols?*

*Show the similarity between an ester and an inorganic salt.*

The formation of esters is known as *esterification*. This is easily reversed. The reverse action is known as *saponification*.

*What must take place in order that a reaction may go to an end?*

*How does the presence of concentrated sulphuric acid aid in the formation of esters?*



PERIODIC CLASSIFICATION OF THE ELEMENTS

Periods	O	I		II		III		IV		V		VI		VII		VIII
	A	B	A	B	A	B	A	B	A	B	A	B	A	B	A	B
First Short	He (4)	Li (7)		Cl (9)	B (11)		C (12)				N (14)	O (16)		F (19)		
Second Short	Ne (20)	Na (23)		Mg (24.3)	Al (27)		Si (28.3)				P (31)	S (32)		Cl (35.5)		
First Long	A (39.9)	K (39.1)		Ca (40)	Sc (44)		Ti (48)			V (51.2)	Cr (52)			Mn (55)		Fe (56)
		Cu (63.6)		Zn (65.4)	Ga (70)		Ge (72.5)			As (75)	Se (79.2)			Br (80)		Co (59)
Second Long	Kr (83)	Rb (85)		Sr (87.6)	Yt (89)		Zr (90.6)			Cb (93.5)	Mo (96)					Ru (102)
		Ag (108)		Cd (112)	In (115)		Sn (119)			Sb (120)	Te (127.5)			I (127)		Rh (103)
Third Long	Xe (130)	Cs (133)		Ba (137)	La (139)		Ce* (140)			Ta (181)	W (184)					Os (191)
		Au (197)		Hg (200)	Tl (204)		Pb (207)			Bi (208)						Ir (193)
Formula of Oxide Formula of Hydride	Nt (222)			Ra (226)			Th (232)				U (238)					
																RO <sub>4</sub>

\* Between cerium and tantalum are fourteen elements which are very similar. Their position in the table is quite odd because they form a series of their own, and, according to their valence, they should belong to the third group.

*elements which are put together in the eighth group*; this seems to take the place of the empty space in the zero group.

(We can show the relation of the elements still better by dividing each group into subgroups, or *families*, A and B. In the first two periods the first three elements are put in Group A, the last three in Group B. The elements which occur in the first half of each long period are placed in Group A; those in the second half in Group B. The zero and eighth groups have no subdivisions.)

The first important fact to be noted about the periodic table is that *all the elements which occur in a group have the same valence toward a given element*. Below the group is written the typical formula for the oxide and for the hydrogen compound of members of that group. The elements in the zero group will not combine and may be considered as having a zero valence.

It will further be seen from the oxides that as we proceed from left to right along any period the valence of the element increases regularly to the seventh. These regularities are very striking; when we remember that this classification of the elements is based solely on the position which they occupy in the series by virtue of their atomic weights, it is even more extraordinary.

In the first two periods (which are short periods) the members of groups V, VI, and VII are active *nonmetals*. In the long periods the three middle elements, which fall in Group VIII, are all metals of a rather inactive sort; at the beginning of the period are active metals, and at the end of the period is an active nonmetal, such as bromine.

In order to see how valuable this periodic table is as a method of classification, let us consider the groups which contain the nonmetals. Group VII B is the halogen family. We have already seen that there is a gradual change in properties as we proceed down the group from fluorine to iodine, the elements becoming less active. Group V B is the nitrogen family, with nitrogen, a typical nonmetal, then phosphorus, arsenic, antimony, and bismuth, a typical metal. As a general rule, when we proceed down a group the elements become less active. In groups IV, V, VI, and VII this loss of activity is accompanied by a gain in metallic properties; the oxides become less acidic and more basic.

The very active metals sodium and potassium we find together in Group I A. The less important elements lithium and rubidium resemble them closely. These metals form oxides which combine with water, giving very strong bases. We have already considered the close resemblance in this group between sodium and potassium. The elements which lie in I B and in Group VIII are also metals, but of a different sort; they are not so active and their hydroxides are weak bases.

In general, we may sum up by saying that *in the periodic system the most active metals are at the top and to the left; the most active nonmetals, to the top and to the right. Certain rather inactive metals lie in the middle of the long period and are found in groups VIII and I B. As we proceed down any group the elements become less active; the nonmetallic elements become more metallic.*

The essential characteristic of an element like sodium is its activity and the formation of a positive ion; this is sometimes called an extreme electropositive character. The essential characteristic of a nonmetal like fluorine is its activity and the formation of a negative ion; this is sometimes called an extreme electronegative character. With these definitions in mind we see that the most electropositive elements are in Group I A, and the most electronegative elements in Group VII B. In between these extremes lie elements which represent the many gradations between these two opposite characteristics.

The periodic system is one of the greatest generalizations in chemistry. While there are apparently numerous inconsistencies in special cases (as in the position of argon and potassium, also in the position of iodine and tellurium), yet on the whole it represents a very orderly arrangement and classification of a multitude of chemical and physical facts.

The table is of value *first* as a *means of classifying* what would otherwise be an almost hopeless and meaningless wilderness of chemical facts. In this way the system is useful not only to the student but also to the chemist and to the scientific investigator. For the latter it points the way to a more fundamental explanation of the whole science of chemistry.

The *second* great use of the system has been in the *prediction of the discovery of new elements*. When Mendelejeff first proposed his table,

he was obliged to leave in it a number of blanks which he believed represented elements not yet discovered. On the basis of the position of these blanks he was able to predict with great accuracy not only the existence of these elements but their physical and chemical properties as well. His prediction was verified in three instances. The following table shows one of these; he prophesied the properties of what he called eka-aluminium, which when it was finally discovered was called *gallium*. There is hardly a scientific prediction in history that is more startling than this.

PROPERTIES OF GALLIUM	PREDICTED	FOUND
Atomic Weight	about 69	69.0
Melting Point	low	30.1°
Specific Gravity	about 5.9	5.93
Action of Air	none	slightly oxidized at red heat
Action on Water	decomposes at red heat	decomposes at high temperatures

Finally, the system has been of use in the *revision of certain atomic weights*. A number of elements were thought at first sight not to fit into the table because of their atomic weights. In all but two or three instances a revision of the atomic weights showed that the earlier values were in error. There are still one or two elements whose atomic weights do not fit them into their proper place in the table. However, very recent developments have shown that the position of these elements is correct.

The position of hydrogen in the system has been a matter of some discussion, but it is not of much consequence. It seems to be rather an odd element. Perhaps the best place for it is in Group I A as it forms a positive ion. The rare gases fit very well into the zero group, and since they have no chemical activity, we do not have to concern ourselves with their valence.

#### QUESTIONS FOR STUDY

1. What does the phrase "periodic function" mean?
2. What natural processes might be called periodic?
3. Sulphur and oxygen both belong to the same group; in what respects are they similar?



4. Illustrate the periodic classification by two periods.
5. Explain how Group O (the inert gases) fit into the periodic system.

#### Experiment 48. Fermentation.

**Apparatus.** Acid bottle, capacity about  $2\frac{1}{2}$  liters; two wide-mouthed bottles, capacity about 500 c. c.; stoppers and tubes as shown in figure; test tube; boiling flask, 250 c. c.; U-tube, 6"; battery jar, 5" high; stoppers and tubes as shown in figure 74; thermometer; watch glass; beaker, 100 c. c.; ring stand with 1 ring and a small clamp; Bunsen burner; wire gauze with asbestos center.

**Material.** Molasses; two yeast cakes; limewater; iodine; quicklime; sodium hydroxide solution; candle or splinter.

##### a. Fermentation.

Arrange apparatus as shown in figure. One set of apparatus will furnish enough fermented liquid for ten pupils to use in part b. Pour one volume of molasses dissolved in six volumes of water into the acid bottle. Break two yeast cakes into small fragments and stir them into a little lukewarm water, so as to form a thin paste.

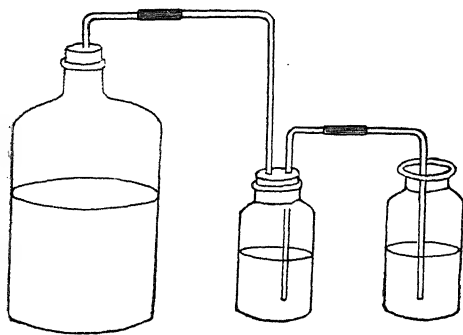


Fig. 74

into the bottle containing the solution of molasses, then shake the bottle so as to distribute the yeast through the molasses. After replacing the stopper in the acid bottle, remove the stopper from the wide-mouthed bottle connected with the acid bottle, and fill the bottle with water. Replace the stopper. Some of the water will be forced into the third bottle, but no

siphoning can take place since air cannot enter the acid bottle.

Allow the apparatus and contents to remain in a warm place for several days (about a week). From time to time, examine the contents of the bottle.

*What change do you notice in the color of the liquid in the acid bottle?*

*What has happened to the water in the wide-mouthed bottle?*

## THE PERIODIC SYSTEM



While the action is in progress, remove the third bottle for a moment and replace it with a test tube containing a little limewater so arranged that the end of the delivery tube will be in the limewater.

*Account for the result.*

After the action has been going on for several days, remove the stopper from the bottle filled with gas and lower into it a lighted candle, or burning splinter.

*Result?*

*What is the name of the gas?*

*b. Fractional distillation.*

Arrange apparatus as shown in figure 75.

Siphon off enough of the fermented liquid to half fill a distilling flask and clamp the flask in place. Insert the thermometer so that its bulb is in the neck of the flask just below the delivery tube.

After connecting the condensing tube to the flask, bring the liquid in the flask to a gentle boil, then regulate the temperature very carefully, in order to prevent froth from being formed and passing over into the distillation tube.

*Why should the water in the battery jar be changed occasionally or lumps of ice be put in the water?*

*At what temperature does the liquid commence to distill over?*

When the temperature has reached 85°, take away the flame. Remove the U-tube from the water, wipe it dry on the outside, and pour enough of its contents into a watch glass to make a circle as big as a two-dime piece. Hold a Bunsen flame to the watch glass for a moment.

*What happens?*

*What is shown to be present by this test?*

*After the action is over, is any liquid left in the watch glass?*

Stop the distillation at 99° C.

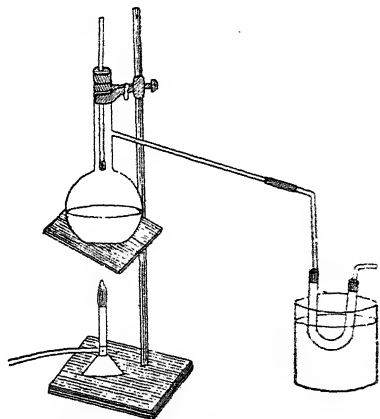


Fig. 75

At the close of the distillation try to ignite a few drops of the distillate.

*Result?*

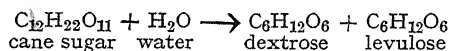
Throw away the liquid in the flask.

Place in the distilling flask a number of small lumps of quicklime. Pour the first distillate back into the flask and redistill it, noticing the temperature of the vapor.

*Why is the lime used?*

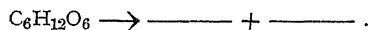
Identify the second distillate (*a*) by its odor, (*b*) by bringing a lighted match to a small portion in a watch glass, and (*c*) by applying the iodoform test to another portion. To make the iodoform test, add to the liquid to be tested a little sodium hydroxide solution, and then iodine, a crystal at a time, and warm. Stop the addition of the iodine before a permanent brown color is obtained. The iodoform will separate as a yellow precipitate and can be recognized by its odor.

During the fermentation, the cane sugar in the molasses was converted into a mixture of two simpler sugars, dextrose and levulose:



This change was brought about by a ferment (invertase) produced during the growth of the yeast plant. Then another ferment (zymase), also from the yeast, caused the change of the simpler sugars into ethyl alcohol ( $\text{C}_2\text{H}_5\text{OH}$ ), and the gaseous product you tested.

*Complete the equation for this change:*



Ethyl alcohol boils at  $78^\circ \text{C}$ .

*How do you explain the rise in the boiling point during the distillation?*



temporarily built into tissues which later undergo oxidation. In this process most of the carbon is finally changed into carbon dioxide and exhaled from the lungs, while the hydrogen is changed into water. The nitrogen is excreted largely in the form of urea  $[\text{CO}(\text{NH}_2)_2]$ .

Broadly speaking, foods may be regarded as fuel from the oxidation of which in the body the energy necessary for the bodily requirements is set free. In the study of foods it is convenient, therefore, to use their *fuel values* (heats of combustion) as a basis of comparison. These values are determined in the calorimeter and expressed in large calories (Cal.), which are 1,000 times as large as the small calorie (cal.).

Now experiments show that the body obtains from each of the three classes of food, when absorbed in the system and oxidized, approximately the following fuel values:

CLASS OF FOODS	CALORIES PER GRAM	CALORIES PER POUND
Carbohydrates	4	1,815
Fats	9	4,082
Proteins	4	1,815

Many studies have been made in order to find out just how much and what kind of food is best adapted for the preservation of health. Evidently many conditions, such as one's age, weight, and occupation, and the climate in which one lives, enters into the problem.

Since the fats and carbohydrates have nearly the same function, it is sufficient in stating food requirements for a period of, say, twenty-four hours, to give simply the weight of protein necessary, together with the total fuel value. The difference between the total fuel value and that of the required protein gives the number of calories to be supplied from fats and carbohydrates. Such a mixture of these two food materials is selected as will suit the taste, provided only that the fuel value of the mixture together with the fuel value of the protein equals the total fuel value required. The following dietary standards proposed by Atwater are generally accepted. They give the food requirements for a period of twenty-four hours.

CHARACTER OF INDIVIDUAL	PROTEINS REQUIRED	FUEL VALUE REQUIRED
Man with very hard muscular work (wood-chopper, football player)	175 g.	5,500 Cal.
Man with moderately active muscular work	125 g.	3,400 Cal.
Man with light to moderate muscular work	112 g.	3,050 Cal.
Man at sedentary, or woman with moderately active, work	100 g.	2,700 Cal.
Man at rest, or woman with light muscular work	90 g.	2,450 Cal.
Boy 15 to 16 years	108 g.	3,060 Cal.
Boy 13 to 14 years or girl 15 to 16	100 g.	2,720 Cal.
Boy 12 to 13 years or girl 14 to 15	87 g.	2,380 Cal.
Boy 10 to 11 years or girl 10 to 12	75 g.	2,040 Cal.
Boy 6 to 9 years	62 g.	1,700 Cal.

It will be noted that the food requirements for the body are stated simply in terms of protein, carbohydrates, and fats. It is very evident that the cost of, say, 100 g. of protein will vary according to the source of the protein. For example, protein obtained in the form of tenderloin beefsteak will cost over eight times as much as an equal weight of protein obtained from dried beans. Again, there is as much nutriment in 1 lb. of wheat flour (cost, about 4 cents) as in  $3\frac{1}{2}$  qt. of oysters (cost, about \$1.40). It is interesting that 4 cents' worth of corn meal or oatmeal contains sufficient protein and total fuel value to supply the needs of the body for twenty-four hours. A diet of this kind, however, is not agreeable, since we demand food that is more highly palatable and more attractive to the senses. By finding the selling prices of the various foods, one can easily determine the relative cost of various food materials obtained from different sources.

Plants build up their tissues from water, nitrogenous matter, and mineral salts absorbed from the soil, and carbon dioxide absorbed from the air. In this process a large amount of energy derived from the sun's rays is absorbed.

Animals, on the other hand, live upon these complex vegetable substances, and in the course of their assimilation the processes which took place in their formation in the plant are largely reversed. The carbon and hydrogen are again liberated, principally as carbon dioxide and water, the nitrogen is secreted largely in the form of urea  $[\text{CO}(\text{NH}_2)_2]$ , while the energy stored up in the plant is utilized as a source of heat and for muscular work. In general, it may be said, therefore, that the

processes involved in plant and animal nutrition are opposite in character; *the plant builds up complex tissues out of simple compounds with energy absorption; the animal decomposes these tissues into simple compounds with evolution of energy.*

#### QUESTIONS FOR STUDY

1. Is the ash obtained in burning a food present in that form in the food?
2. Why do different nations use different classes of foods?
3. Mention some foods that are rich in water; in protein; in fats; in carbohydrates.
4. Find out the current prices of common foods, then work out a list of the most economic ones to use as a source of protein; of carbohydrates; of fats.

#### **Experiment 49. Food Constituents—Fats, Proteins.**

**Apparatus.** Test tubes; Bunsen burner; ring stand with one large and one very small ring; two porcelain evaporating dishes; beaker, 150 c. c.

**Material.** Benzene; filter paper; Geis's biuret reagent, made by adding 3 per cent copper sulphate solution to 10 per cent potassium hydroxide solution a drop at a time, until a faint but perceptible blue color is imparted to the resulting solution; foods for testing, such as fat meat, flour, milk, nuts, sugar, peas, etc.

Test several foods as directed below, and fill in the tabular form.

##### *a. Test for oil or fat.*

Divide as finely as possible the sample to be tested. Use enough to fill a test tube to the depth of one half to three fourths of an inch. Add benzene to cover the solid to the depth of one half inch.

*Caution! Benzene is inflammable!*

Warm the contents gently by standing the test tube in a beaker of hot water. Allow the test tube to stand for five minutes, with occasional shaking and warming.

Fold a filter paper as for a funnel. Suspend it in a small ring in the ring stand. Pour a few drops of the liquid from the test tube into the point of the filter paper, and allow the benzene to evaporate completely. If the food contains oil or fat, more or less of a line of grease will be left around the stained spot on the paper. This can be more easily





sulphuric acid in such a manner that it slides down the side of the tube and makes a separate layer below the mixture already present in the tube. A purple ring will form between the two layers if the original solution contained any carbohydrate.

*Result?*

*b. Test for starch.*

To a pinch of starch add 5 c. c. of water and boil the mixture in a test tube till a gelatinous mass has been formed. Dilute this by nearly filling the tube with water. Add a drop of iodine solution.

*Result?*

*c. Test for glucose.*

Dissolve a little glucose in 10 c. c. of water. Heat in a second test tube 10 c. c. of Fehling's solution to the boiling point. To this hot liquid add a little of the solution of glucose.

*Result?*

*d. Formation of glucose from starch and cane sugar.*

To about 1 gram add 10 c. c. of water and 3 or 4 drops of hydrochloric acid. Boil the mixture for 10 minutes. Test the resulting solution for glucose.

*Result?*

In a similar way, boil cane sugar solution with hydrochloric acid and test the product for glucose.

*Result?*

*Under what circumstances can Fehling's solution be used as a test for cane sugar?*

*e. Test of samples.*

Test the samples of food given you, first for carbohydrates in general, and then for starch and glucose. If a food contains neither of these, test it for cane sugar. Use the food in a finely divided state, and boil about 1 gram of it with 10 c. c. of water for five minutes before making the test.

*Record your results in the table.*

*Why cannot the test for cane sugar be used in the presence of starch?*

# FOODS

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TABLE

FOOD TESTED	STARCH PRESENT OR ABSENT ?	GLUCOSE (OR CANE SUGAR) PRESENT OR ABSENT ?







TABLE I. PHYSICAL CONSTANTS OF THE IMPORTANT ELEMENTS

ELEMENT	SYMBOL	ATOMIC WEIGHTS		VALENCE	SPECIFIC GRAVITY		MELTING POINT °C.	BOILING POINT °C.
		APPROXI- MATE	EXACT O=16		WATER = 1	AIR=1		
Aluminium	Al	27	27.1	III	2.7	—	657	2200
Antimony	Sb	120	120.2	III V	6.6	—	630	1600
Argon	A	40	39.9	—	—	1.38	-118	-186
Arsenic	As	75	74.96	III V	5.7	—	—	450 volatile
Barium	Ba	137	137.37	II	3.8	—	850	950
Bismuth	Bi	208	208.0	III V	9.7	—	269	1435
Boron	B	11	10.9	III	2.4	—	—	3500
Bromine	Br	80	79.92	I	3.1	—	-7.3	59
Cadmium	Cd	112	112.4	II	8.6	—	321	778
Calcium	Ca	40	40.07	II	1.8	—	805	—
Carbon	C	12	12.00	IV	amor- phous 1.4 to 1.9	—	infusible	3500
Chlorine	Cl	35.5	35.46	I	—	2.49	-102	-33.6
Chromium	Cr	52	52.0	II III VI	6.9	—	1505	2200
Cobalt	Co	59	58.97	II	8.7	—	1490	—
Copper	Cu	63.6	63.57	I II	8.9	—	1083	2310
Fluorine	F	19	19.0	I	—	1.26	-223	-187
Gold	Au	197	197.2	I III	19.3	—	1062	2350
Helium	He	4	4.00	—	—	0.137	-271	268.5
Hydrogen	H	1	1.008	I	—	0.07	-259	-252
Iodine	I	127	126.92	I	4.9	—	114	184
Iron	Fe	56	55.84	II III	7.8	—	1520	2450
Lead	Pb	207	207.20	II IV	11.3	—	327	1525
Lithium	Li	7	6.94	I	0.59	—	186	1400
Magnesium	Mg	24	24.32	II	1.7	—	650	1120
Manganese	Mn	55	54.93	II IV	7.4	—	1225	1900
Mercury	Hg	200	200.6	I II	13.6	—	-38.8	357
Neon	Ne	—	20.2	—	—	—	-253	-238
Nickel	Ni	58.7	58.68	II	8.7	—	1450	—
Nitron	Nt	—	222.4	—	—	—	—	—
Nitrogen	N	14	14.008	III V	—	0.96	-213	-195
Oxygen	O	16	16.00	II	—	1.10	-227	-182
Phosphorus	P	31	31.04	III V	1.8	—	44.1	290
Platinum	Pt	195	195.2	IV	21.1	—	1755	—
Potassium	K	39	39.10	I	0.87	—	62.5	712
Silicon	Si	28	28.3	IV	2.4	—	1420	3500
Silver	Ag	108	107.88	I	10.5	—	961	1955
Sodium	Na	23	23.0	I	0.97	—	97.6	750
Strontium	Sr	87	87.63	II	2.5	—	900	—
Sulphur	S	32	32.06	II IV VI	2.0	—	114.5	444.6
Tin	Sn	119	118.7	II IV	7.0-7.3	—	232	1525
Zinc	Zn	65	65.37	II	7.1	—	419	918

TABLE II. PHYSICAL CONSTANTS OF COMMON GASES

	FORMULA	MOLECULAR WEIGHT	DENSITY	SPECIFIC GRAVITY	SOLUBILITY IN 1 VOLUME OF WATER
Acetylene	$C_2H_2$	26.0	1.17	.906	—
Air	—	—	1.283	1.00	—
Ammonia	$NH_3$	17.0	.771	.596	72722
Argon	A	39.9	1.78	1.79	—
Carbon Dioxide	$CO_2$	44.0	1.98	1.53	100
Carbon Monoxide	CO	28.0	1.25	.967	—
Chlorine	$Cl_2$	70.92	3.22	2.49	150
Helium	He	4	.178	.137	—
Hydrogen	$H_2$	2.016	.0896	.067	2
Hydrogen Chloride	HCl	36.45	1.64	1.27	46400
Hydrogen Sulphide	$H_2S$	34.0	1.54	1.19	250
Methane	$CH_4$	16.0	.717	.554	5
Nitric Oxide	NO	30.0	1.34	1.37	7
Nitrogen	$N_2$	28.02	1.26	.97	2
Nitrous Oxide	$N_2O$	44.0	1.98	1.53	130
Oxygen	$O_2$	32.00	1.429	1.105	5
Sulphur Dioxide	$SO_2$	64.1	2.93	2.26	4400

The values above are experimentally determined. Density is given in grams per liter. Specific gravity is based on the air standard. Solubility=volumes of gas dissolved by 100 volumes of water.

TABLE III. TABLE OF SOLUBILITIES

	Aluminium	Ammonium	Antimony	Arsenious	Barium	Bismuth	Cadmium	Calcium	Cobalt	Copper	Ferrous (Fe++)	Ferric (Fe+++)	Lead	Magnesium	Manganese	Mercurous (Hg+)	Mercuric (Hg++)	Nickel	Potassium	Sodium	Silver	Zinc
Acetate	S	S	S	S	S	S	S	S	S	S	S	—	S	S	S	S	S	S	S	S	S	S
Bromide	S	S	P	S	S	P	S	S	S	S	S	S	P	S	S	I	S	S	S	S	Ia	S
Carbonate	—	S	—	—	I	I	I	I	I	I	I	—	I	I	I	I	I	I	S	S	I	I
Chlorate	S	S	—	—	S	—	S	S	S	S	—	—	S	S	—	S	S	—	S	S	S	S
Chloride	S	S	P	S	S	P	S	S	S	S	S	S	P	S	S	I	S	S	S	S	Ia	S
Chromate	—	S	—	—	I	I	—	I	I	I	—	—	Ia	S	—	P	P	—	S	S	I	—
Hydroxide	I	S	P	S	S	I	I	P	I	I	I	I	I	I	I	—	I	I	S	S	—	I
Iodide	S	S	P	P	P	I	S	S	S	—	S	—	I	S	S	I	I	S	S	S	Ia	S
Nitrate	S	S	—	—	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S
Oxide	I	—	I	P	S	I	I	P	I	I	I	I	I	I	I	I	I	I	S	S	I	I
Phosphate	I	S	—	—	I	I	I	I	I	I	I	I	I	I	P	I	I	I	S	S	I	I
Sulphate	S	S	—	—	Ia	S	S	P	S	S	S	S	Ia	S	S	P	S	S	S	S	P	S
Sulphide	—	S	Ia	Ia	S	Ia	Ia	P	I	Ia	P	—	Ia	S	I	Ia	Ia	I	S	S	Ia	I
Sulphite	—	S	—	—	I	—	S	P	I	S	P	—	I	P	—	—	—	I	S	S	S	P

S=soluble in water. I=insoluble in water. P=slightly (partly) soluble in water.  
Ia=insoluble in water and dilute acids.



TABLE IV. GENERAL RULES FOR SOLUBILITY

Certain generalizations can be made concerning compounds shown in Table III. The exceptions to these generalizations are few and unimportant.

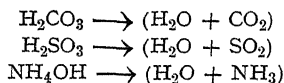
1. All *sodium, potassium* and *ammonium* compounds are *soluble* in water.
2. All *nitrates, chlorates, and acetates* are *soluble* in water.
3. All *chlorides* are *soluble*, except those of silver, mercury (mercurous), and lead (lead slightly soluble).
4. All *sulphates* are *soluble*, except those of lead, barium, strontium, and calcium (calcium slightly soluble). The silver and the mercurous sulphates are only moderately soluble.
5. All *carbonates* are *insoluble*, except those of sodium, potassium, and ammonium.
6. All *oxides* and *hydroxides* are *insoluble*, except those of ammonium, sodium, potassium, and barium; calcium hydroxide is slightly soluble.

TABLE V. VOLATILITY OF COMPOUNDS THAT MAY RESULT FROM DOUBLE DECOMPOSITIONS

1. Compounds volatile at ordinary temperatures :



2. Compounds decomposing at ordinary temperatures yielding volatile products:-



3. Compounds volatile at varying temperatures below 338° (boiling point of sulphuric acid):

Boiling Point		Boiling Point °	
H <sub>2</sub> O	100°	HNO <sub>3</sub>	86°
HCl (aqueous solution)	110°	HNO <sub>3</sub> (aqueous solution)	120°
HBr (aqueous solution)	126°	HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub>	118°

TABLE VI. SOLUBILITY OF SOLIDS

Column A gives the number of grams of the anhydrous salt which will dissolve in 100 grams of water at 15° C.

Column B gives the number of grams of the anhydrous salt in 100 grams of water at the boiling point of the saturated solution, given in Column C.

	A	B	C
Ammonium Chloride	35	87	115°
Barium Chloride	35	54	104°
Calcium Chloride	65	178	150°
Calcium Hydroxide	0.17	.05	—
Copper Sulphate	20	60	102°
Potassium Chloride	33	57	108°
Potassium Nitrate	26	335	115°
Sodium Chloride	36	41	109°
Sodium Sulphate	14	43	103°
Sugar	196	—	—

TABLE VII. IMPORTANT IONS CONTAINING MORE THAN ONE ELEMENT

Acetate ( $\text{C}_2\text{H}_3\text{O}_2$ ) <sup>-</sup>	Ferrocyanide [ $\text{Fe}(\text{CN})_6$ ] <sup>----</sup>
Ammonium ( $\text{NH}_4$ ) <sup>+</sup>	Ferricyanide [ $\text{Fe}(\text{CN})_6$ ] <sup>----</sup>
Bicarbonate ( $\text{HCO}_3$ ) <sup>-</sup>	Hydroxide ( $\text{OH}$ ) <sup>-</sup>
Biborate ( $\text{B}_4\text{O}_7$ ) <sup>--</sup>	Nitrate ( $\text{NO}_3$ ) <sup>-</sup>
Bisulphate ( $\text{HSO}_4$ ) <sup>-</sup>	Nitrite ( $\text{NO}_2$ ) <sup>-</sup>
Borate ( $\text{BO}_3$ ) <sup>---</sup>	Oxalate ( $\text{C}_2\text{O}_4$ ) <sup>--</sup>
Carbonate ( $\text{CO}_3$ ) <sup>--</sup>	Phosphate ( $\text{PO}_3$ ) <sup>---</sup>
Chlorate ( $\text{ClO}_3$ ) <sup>-</sup>	Sulphate ( $\text{SO}_4$ ) <sup>--</sup>
Chromate ( $\text{CrO}_4$ ) <sup>--</sup>	Sulphite ( $\text{SO}_3$ ) <sup>--</sup>
Cyanide ( $\text{CN}$ ) <sup>-</sup>	Tartrate ( $\text{C}_4\text{H}_4\text{O}_6$ ) <sup>--</sup>

TABLE VIII. DEGREE OF IONIZATION OF IMPORTANT COMPOUNDS

Showing the fractional part of the solute which is ionized at the concentration given, and at normal temperatures :

SOLUTE	CONCENTRATION	IONIZATION
Nitric Acid	normal	0.82
Hydrochloric Acid	normal	0.78
Hydrochloric Acid	one-tenth normal	0.87
Sulphuric Acid	normal	0.51
Oxalic Acid	one-tenth normal	0.50
Phosphoric Acid	half normal	0.17
Tartaric Acid	one-tenth normal	0.08
Acetic Acid	one-tenth normal	0.004
Carbonic Acid	—	0.002
Hydrogen Sulphide	—	0.001
Boric Acid	—	0.0001
Potassium Hydroxide	normal	0.77
Sodium Hydroxide	normal	0.73
Ammonium Hydroxide	—	0.004
Barium Hydroxide	normal	0.69
Barium Hydroxide	one-tenth normal	0.92
Strontium Hydroxide	one-tenth normal	0.93
Calcium Hydroxide	one-tenth normal	0.50
Salts <sup>1</sup> of Type $R^+A^-$	one-tenth	0.85
Salts of Type $R_2^{++}A^{--}$	one-tenth	0.72
Salts of Type $R^{++}A_2^{--}$	one-tenth	0.72
Salts of Type $R^{++}A^{--}$	one-tenth	0.50

<sup>1</sup> R stands for any metal having the valence indicated ; A stands for a negative ion with indicated valence.

TABLE IX. THE METRIC SYSTEM AND ITS EQUIVALENTS

## A. FUNDAMENTAL UNITS

The *International Standard Meter* is the distance between two lines, at 0° Centigrade, on a platinum-iridium bar deposited at the International Bureau of Weights and Measures near Paris, France.

## MEASURES OF LENGTH

10 millimeters (mm.)	= 1 centimeter (cm.)
10 centimeters	= 1 decimeter (dm.)
10 decimeters	= 1 meter (m.)

The *International Standard Kilogram* is the weight of a mass of platinum-iridium deposited at the International Bureau of Weights and Measures.

The *liter* is practically equal to a cubic decimeter and it is measured by the quantity of distilled water which, at its maximum density (4° centigrade), will counterpoise the standard kilogram.

Since a *liter* contains 1,000 cubic centimeters, one cubic centimeter of water, at 4° centigrade, weighs 1 gram.

## B. IMPORTANT METRIC RELATIONS

MEASURE	RELATION
Linear	Millimeter = 0.001 meter Centimeter = 0.01 meter Decimeter = 0.1 meter
Capacity	Cubic centimeter = 0.001 liter (practically)
Weight	Milligram = 0.001 gram Centigram = 0.01 gram Decigram = 0.1 gram Kilogram = 1,000 grams

### C. COMPARISON OF METRIC WITH OTHER WEIGHTS AND MEASURES

METRIC UNIT	EQUIVALENT		MEASURE
	APPROXIMATE	EXACT	
Meter	—	39.37 inches	U. S. Linear
Centimeter	0.4 inch	0.3937 inch	U. S. Linear
Liter	1.06 quarts	1.05668 quarts	U. S. Liquid
Kilogram	2.2 pounds	2.20462 pounds	U. S. Avoirdupois
Gram	15.4 grains	15.43235 grains	U. S. Avoirdupois

UNIT	U. S. MEASURE	EQUIVALENT	
		APPROXIMATE	EXACT
Yard	Linear	0.9 meter	0.914402 meter
Inch	Linear	2.5 centimeters	2.54001 centimeters
Gallon	Liquid	3.8 liters	3.78543 liters
Quart	Liquid	0.95 liter	0.94636 liter
Fluid Ounce	Liquid	29.6 c. c.	29.574 c. c.
Pound	Avoirdupois	0.45 kilogram	0.45359 kilogram
Pound	Avoirdupois	453.6 grams	453.59 grams
Ounce	Avoirdupois	28.3 grams	28.3495 grams
Grain	Avoirdupois	0.06 gram	0.0648 gram

TABLE X. PRESSURE OF WATER VAPOR, OR AQUEOUS TENSION

IN MILLIMETERS OF MERCURY

TEMPERATURE, ° C.	PRESSURE, MM.	TEMPERATURE, ° C.	PRESSURE, MM.
0.0	4.6	21.5	19.1
5	6.5	22	19.7
10	9.2	22.5	20.3
10.5	9.5	23	20.9
11	9.8	23.5	21.5
11.5	10.1	24	22.1
12	10.5	24.5	22.8
12.5	10.8	25	23.5
13	11.2	25.5	24.2
13.5	11.5	26	25.0
14	11.9	26.5	25.7
14.5	12.3	27	26.5
15	12.7	27.5	27.3
15.5	13.1	28	28.1
16	13.5	28.5	28.9
16.5	14.0	29	29.8
17	14.4	29.5	30.7
17.5	14.9	30	31.6
18	15.4	40	54.9
18.5	15.9	50	92.1
19	16.4	60	149.2
19.5	16.9	70	233.8
20	17.4	80	355.4
20.5	17.9	90	526.0
21	18.5	100	760.0



## CHEMICAL LAWS AND THEORIES

*Boyle's Law.* The volume of any gas varies inversely as the pressure, if the temperature remains the same.

*Charles's Law.* The volume of any gas varies directly as the absolute temperature, if the pressure remains the same.

*Gay-Lussac's Law (Law of Volumes).* The relative combining volumes of gases and the volume of the product, if gaseous, may be expressed by small whole numbers.

*Avogadro's Hypothesis.* Equal volumes of gases under like condition of temperature and pressure contain the same number of molecules.

*Law of Definite Proportions.* Every compound has a definite composition by weight.

*Law of Multiple Proportions.* In a series of compounds that are made up of the same elements, a simple ratio exists between the weights of any one element that combine with a fixed weight of another element.

*Law of Conservation of Mass.* There is no change in the total quantity (mass) of matter taking part in any chemical change.

*Atomic Hypothesis.* Formulated by John Dalton. Made up of four assumptions:

1. That matter is made up of small particles, called atoms.
2. These atoms possess the power of attracting or holding on to other atoms.
3. That atoms do not divide in taking part in chemical changes.
4. That all the atoms of an element are exactly alike, but differ from atoms of every other element.

*Theory of Ionization.* Acids, bases, and salts, when dissolved in water, decompose into particles charged with electricity.





## LIST OF SUPPLIES

The following list is an estimate of the material advisable to purchase for a class of ten pupils, provided each student performs all of the experiments described. It names a generous supply of apparatus and chemicals, allowing for a reasonable amount of breakage.

### GENERAL APPARATUS

- 2 balances, platform, with weights for weighing from 0.1 gram to 1,000 grams.
- 10 balances, horn pan,  $7\frac{1}{2}$ " beam, with weights for weighing from 0.01 gram to 100 grams.
- 1 barometer.
- 10 blowpipes, 8".
- 10 brushes, test tube.
- 10 burners, Bunsen, or 10 alcohol lamps.
- 10 candles.
- 10 clamps, iron, small, for burettes, etc.
- 1 gross corks, assorted sizes.
- 1 set cork borers, 6 in. set.
- 12 crucibles, porcelain, with lids, No. 00.
- 12 porcelain evaporating dishes.
- 10 droppers, medicine.
- 10 files, triangular, 5".
- 10 pr. forceps, iron, 4".
- 10 squares gauze, iron wire with asbestos center,  $5" \times 5"$ .
- 10 test-tube holders.
- 1 magnifying glass.
- 5 mortars with pestles,  $3\frac{1}{2}"$ .
- 10 pkg. filter paper, qualitative, good quality, 4".
- 10 pinch cocks, Mohr's, medium.
- 10 racks, test-tube, for 12 tubes.
- 1 pair shears.
- 1 sheet sandpaper.
- 5 spatulas, horn, 6".
- 10 spoons, deflagrating, diameter of bowl 1 cm.
- 10 stands, iron, ring, 3 rings.
- 2 lbs. stoppers, rubber, assorted sizes, Nos. 0-5, 1- and 2-holed.
- 2 thermometers, chemical,  $0^{\circ}$ - $250^{\circ}$  C.
- 1 spool thread, No. 50, cotton.
- 10 triangles, pipestem, size to support No. 00 porcelain crucible.
- 10 troughs, pneumatic, or sinks.
- 25 ft. tubing, rubber, inside diameter  $\frac{3}{16}"$ .

- 10 ft. tubing, rubber, inside diameter  $\frac{3}{8}$ ".
- 1 yardstick.

## GLASSWARE

- 24 beakers, 100 c. c.
- 24 beakers, 150 c. c.
- 24 beakers, 250 c. c.
- 10 beakers, 500 c. c.
- 40 bottles, reagent, 4 oz.
- 30 bottles, wide-mouthed, 4 oz.
- 40 bottles, wide-mouthed, 8 oz. with 2-holed rubber stoppers to fit.
- 2 bottles, wide-mouthed, 16 oz. with 2-holed rubber stoppers to fit.
- 10 bottles, acid,  $2\frac{1}{2}$  liters.
- \*20 burettes, 50 c. c., graduated to  $\frac{1}{10}$  c. c.
- 1 condenser, Liebig, 15".
- \*12 flasks, distilling, 250 c. c.
- 12 flasks, Florence, 50 c. c.
- 12 flasks, Florence, 250 c. c., with 2-holed rubber stoppers to fit.
- \*12 flasks, Erlenmeyer, 250 c. c., with 2-holed rubber stoppers to fit.
- 10 graduates, graduation marks to 1 c. c.
- 1 graduate, 1,000 c. c., graduation marks to 10 c. c.
- 10 jars, battery, about 4"  $\times$  5".
- 40 glass plates, 4"  $\times$  4".
- 40 plates, cobalt glass, for flame tests, 3"  $\times$  2".
- 1 retort, tubulated, with ground glass stopper (for teacher's use).
- 1 lb. glass rods, 3 mm. diameter.
- 10 tubes, gas measuring, 50 c. c., graduated to  $\frac{1}{10}$  c. c.
- 15 tubes, test, hard glass, for ignition, 6"  $\times$   $\frac{5}{8}$ ", with one-holed cork stoppers to fit.
- \*12 tubes, test, side arm, 6".
- 15 dozen tubes, test, soft glass, medium walls, for heating, 6"  $\times$   $\frac{3}{4}$ ".
- 2 dozen tubes, test, soft glass, medium walls, for heating, 4"  $\times$   $\frac{1}{2}$ ".
- 15 tubes, thistle, 10", stem,  $\frac{3}{16}$ " in diameter.
- 12 tubes, U, 4", with one-holed rubber stoppers to fit.
- 2 lb. tubing, soft glass, medium walls for bending, outside diameter 4 mm.
- 24 watch glasses, diameter 3".

## CHEMICALS AND OTHER SUPPLIES

- 1 lb. acid, acetic, 30% c. p.
- 2 oz. acid, boric, c. p.
- 10 lbs. acid, hydrochloric, c. p., sp. gr., 1.19.
- 6 lbs. acid, nitric, c. p., sp. gr., 1.42.
- $\frac{1}{2}$  lb. acid, oxalic, crys., c. p.
- 9 lbs. acid, sulphuric, c. p., sp. gr., 1.84.
- 2 qts. alcohol, ethyl, 95%.

- 1 pt. alcohol, methyl (wood alcohol).
- 1 oz. alpha naphthol.
- 8 oz. aluminium sulphate, cryst.
- 1 lb. ammonium chloride, c. p.
- 10 lbs. ammonium hydroxide, c. p., sp. gr., 0.9.
- 8 oz. ammonium nitrate, cryst. c. p.
- 1 oz. ammonium oxalate, cryst. c. p.
- 4 oz. ammonium sulphate, c. p.
- 1 lb. barium chloride, cryst. c. p.
- 4 oz. barium nitrate, c. p.
- 2 oz. benzol.
- 8 oz. bleaching powder.
- 1 lb. boneblack.
- 1 oz. bromine.
- $\frac{1}{2}$  lb. calcite.
- 5 lbs. calcium carbonate, marble chips.
- 1 lb. calcium chloride, granular, for drying tubes.
- 2 oz. calcium nitrate, c. p.
- 2 lbs. calcium oxide, good quality of lime in tin can.
- 5 lbs. calcium sulphate, plaster of Paris, fine.
- 1 lb. carbon disulphide.
- 4 oz. carbon tetrachloride.
- $\frac{1}{2}$  lb. chalk, precipitated.
- 1 oz. charcoal, wood, powdered.
- $\frac{1}{2}$  lb. chloroform.
- 2 oz. chromium sulphate, c. p. (or chrome alum).
- $\frac{1}{2}$  yd. cloth, bleachable color, for bleaching with chlorine.
- 1 yd. cloth, cotton, bleached.
- $\frac{1}{8}$  yd. cloth, woolen.
- 1 oz. cobalt nitrate, cryst., c. p.
- 4 oz. copper foil,  $\frac{1}{16}$ " thick.
- $\frac{1}{2}$  lb. copper sheet,  $\frac{1}{16}$ " thick.
- 2 lbs. copper turnings, clean, fine.
- $\frac{1}{2}$  lb. spool copper wire, No. 30.
- 1 oz. copper oxide, powdered, c. p.
- 1 oz. copper oxide, wire form.
- 1 oz. copper sulphate, anhydrous.
- 1 lb. copper sulphate, cryst., c. p.
- 1 oz. cotton, absorbent.
- 2 oz. Fehling's solution, two solutions in separate bottles.
- 8 oz. glucose.
- 1 lb. hydrogen peroxide.
- 1 oz. iodine, resublimed.
- 4 oz. iron chloride, ferric, c. p.
- 2 oz. iron filings, fine, clean.
- 1 lb. iron sulphate, cryst., c. p.

- 4 oz. iron sulphide, ferrous, in sticks, for  $H_2S$ .
- 1 lb. spool iron wire, No. 16.
- 2 oz. lead acetate.
- 1 lb. lead shot, No. 10.
- 1 oz. litmus cubes.
- $\frac{1}{2}$  quire each, red and blue litmus paper.
- 1 oz. magnesium, ribbon.
- 1 lb. magnesium sulphate, cryst.
- 2 lbs. manganese dioxide, fine, granular, free from carbon.
- 4 oz. manganese dioxide, c. p.
- 1 oz. mercuric nitrate.
- 8 oz. mercuric oxide, red.
- 1 oz. mercurous nitrate.
- 1 pt. molasses, good quality.
- 10 grams phenolphthalein.
- $\frac{1}{2}$  oz. phosphorus, red.
- 1 oz. phosphorus, yellow.
- 1 lb. potassium and aluminium sulphate, alum.
- 10 oz. potassium bromide.
- 1 lb. potassium chlorate, cryst., c. p.
- \*2 oz. potassium chromate.
- 1 lb. potassium ferricyanide.
- 2 oz. potassium ferrocyanide.
- 2 lbs. potassium hydroxide, c. p., by alcohol.
- 2 oz. potassium iodide, c. p.
- 2 lbs. potassium nitrate, cryst., c. p.
- 1 oz. potassium permanganate, c. p.
- 1 oz. potassium oxalate.
- 8 oz. potassium sulphate, c. p.
- 1 lb. salt, rock (or crude Chinese salt).
- 4 oz. silver nitrate, c. p.
- 2 oz. sodium.
- 1 lb. sodium carbonate, cryst., washing soda.
- 8 oz. sodium carbonate, pure, dry.
- 5 lbs. sodium chloride, salt, fine.
- 1 oz. sodium chromate, powdered.
- 1 lb. sodium hydroxide, c. p., by alcohol.
- 1 oz. sodium iodide (or potassium iodide).
- 1 lb. sodium nitrate, c. p.
- 2 oz. sodium peroxide.
- 8 oz. sodium phosphate, c. p.
- 2 lbs. sodium sulphate, cryst.
- 1 lb. sodium sulphite, pure, dry.
- 1 lb. sodium tetraborate (borax).
- 1 lb. starch, corn.
- 2 lbs. sulphur, roll.

- 8 oz. tin, granulated.
- 1 qt. vinegar, cider.
- 1 oz. wool, glass, fine Bohemian.
- \*1 pkg. wool, steel, fine.
- 4 oz. yarn, woolen, white.
- $\frac{1}{2}$  lb. zinc, sheet.
- 2 lbs. zinc, granulated (mossy).
- 2 oz. zinc dust.
- 8 oz. zinc sulphate.
- 1 oz. zinc nitrate.

\*Those supplies marked with an asterisk are less necessary.

A very few articles, such as flour, yeast, foods for testing, and other articles of common household use have been omitted from the above list.



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